







# DOMESTIC SCIENCE

*The Science of  
Domestic Economy and Hygiene  
Treated Experimentally*

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WITH NUMEROUS ILLUSTRATIONS

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**ASIATIC SOCIETY  
CALCUTTA.**

**THOMAS NELSON AND SONS**

*London, Edinburgh, and New York*

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1900





## P R E F A C E.

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THE simple course of Experimental Science mapped out in these little manuals covers the subject of Domestic Science as set out in Specific Subject XIV. of the Education Department's Code. The author entreats the teacher who uses these books to insist on the girls working all the experiments with their own hands. His experience with girls has taught him that they soon become adepts at manipulation, and that they are not one whit behind boys in their power of grasping the truths that the experiments are intended to make manifest. Hence he advises, and that most strenuously, that the simple experiments described should be performed by every girl, who will certainly more firmly grasp the fundamentals of Domestic Economy by practically proving with her own hands their reality.

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*January 22, 1900.*



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# DOMESTIC SCIENCE.

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## Part I.

### • I. WEIGHING AND MEASURING.

**The Housewife a Skilled Worker.**—To be able to produce a first-rate article, whether in the kitchen or in the workshop, is the praiseworthy ambition of the good worker. This requires a thorough knowledge both of the material to be worked up and of the tools to be used in the working; hence the good housewife, who wishes to become a skilled worker, must study both what she has to work upon and what she has to work with. She cannot afford to be ignorant of science, since it is only by conforming to the laws of science that she may hope to do her best, whether in cooking, in washing, in ventilating, or in disinfecting. To give such a knowledge of elementary science, by experiment, as will enable the housewife to become a skilled worker, is the aim of this little book.

**Weighing and Measuring.**—We shall begin with some simple exercises in these two operations, with which every housekeeper ought to be familiar. We shall,

first of all, ask what is meant by weighing and measuring, and shall discuss the various weights and measures, in order to show how best to use them. Knowing the origin of the weights and measures, she so frequently uses, the housewife will take a keener interest in her work, which she will also perform with greater care and intelligence.

**The Meaning of Measuring.**—When the draper's assistant hands his customer over the counter a length of calico, which he says is a dozen yards, what has he done, and what does he mean? He has seen how many times a certain stick will divide, or go, into the length of calico, and he means that the calico is twelve times as long as the stick. Here is plainly shown the chief thing needed in measurement of length—namely, a **standard, or unit of measurement**, which in this case is the yard-stick, by referring to which we are able to understand how much calico there is in the length. When, therefore, we know what “**yard**” means, and not before, we know how much twelve yards of calico really represents; and it is evident that in measuring calico, or anything else, we simply **compare it with this unit**, and what the measurer really finds out is the ratio between the stick and the cloth—that is, how many times the former is contained in the latter. And so it is always. In measuring **length, area, volume, or weight**, we simply compare the body to be measured with a standard or unit that has been agreed upon by the parties concerned.

**The Yard.**—We in England have agreed to accept the yard as the standard, or unit, of linear measure; and it is very necessary that a correct definition of the term yard should be given.

In the Standards Office at Westminster is a bar of bronze or gun-metal 38 inches long, and 1 square inch in transverse section. One inch from each end is a plug of gold, across the centre of which is a scratch, and the distance between these two scratches is the standard or imperial yard. Hence, we define the yard as the distance between the centres of two gold plugs in a certain bronze bar kept in the Standards Office.

In fixing the yard, there is nothing to guide us except this bronze bar, so that the standard is a purely arbitrary and not a natural one.

**Origin of the Name Yard.**—The first units of measurement were taken from parts of the body. This is evidently the case with respect to the foot; and when we learn that the old form of the word yard is *gird*, meaning girdle or distance round the waist, it is clear that this is the case with respect to the term yard also. So it is with the terms *span* and *cubit*—the former meaning the length from the end of the thumb to the tip of the little finger when the hand is stretched to its utmost; the latter being the length from the elbow to the end of the middle finger. *Inch*, on the other hand, is the Latin *uncius*, meaning twelfth part. These units were very unscientific, because very variable; hence arose the need for a fixed and invariable standard, such as is the imperial yard.

**Multiples and Submultiples.**—In dealing with lengths very much greater or very much less than a yard, it is very inconvenient to express these lengths in yards: For this reason we have, starting from the yard, a

Fig. 1.—The Standard or Imperial Yard





series of measures, such as the **pole**, the **chain**, the **furlong**, and the **mile**, which expresses the yard multiplied by  $5\frac{1}{2}$ , 22, 220, and 1,760 respectively. These measures are accordingly termed **multiples of the yard**. Another series of measures, the **foot** and the **inch**, expresses the yard divided by 3 and  $3\frac{1}{2}$  respectively. These measures are appropriately termed the **submultiples of the yard**. It will be noticed that the names multiple and submultiple correspond to product and quotient, since multiples are obtained by multiplying and submultiples by dividing the unit by certain numbers.

**Bad and Good Multiples and Submultiples.**—There is no reason why one number should be used in the multiplication and division of the unit rather than another, except that the number so chosen should be readily remembered and easily manipulated. Unfortunately, this has not been done in building up our English system of weights and measures, which, as every school-girl knows to her cost, are hard to remember and difficult to use in calculation.

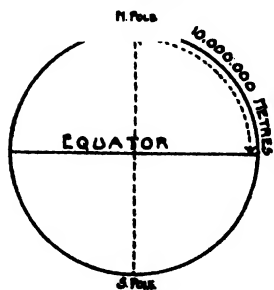


Fig. 2.

**The Metric System.**—The French or metric system of weights and measures is much more convenient. The unit of length is the **metre**, which was intended to be the forty-millionth part of the earth's circumference, or the ten-millionth part of the earth's

quadrant, so that the distance from the equator to the poles would be ten million metres. In conse-

quence of an error, this is not the case, and the metre is best defined as the length between two scratches on a bar of metal deposited in the French Standards Office at Paris. It is not, however, so much in the choice of the unit, as in the method of multiplying and dividing it, that the French system is so convenient. To clearly show how convenient this system really is, we must make a brief reference to our method of writing numbers.

**The Denary or Decimal Scale of Notation.**—If we write down 3·333, each figure increases tenfold in value as we go to the left. There is no natural reason for this selection of ten, except that we have agreed, as a matter of convenience, that it should be the measure of increase or decrease in value according to the position of a figure. Twelve would have been a more convenient base\* from the point of view of duodecimals and the conversion of vulgar fractions into decimals; but ten was probably taken as the foundation of our system of counting, because of the use of man's ten fingers in primitive calculation. To return to our example. We notice that if a certain figure 3 is removed three places from the unit to the left it represents  $3 \times 10^3$ —that is, three times three tens multiplied together, or  $3 \times 10 \times 10 \times 10$  and so always the figure or integer multiplies the product of as many tens as it counts places to the left of the unit. In the year 1562, Simeon Stephens of Bruges, in Flanders, advanced a further step in utiliz-

°	Decimally	$\frac{1}{2} = \cdot 333$	.....	Duodecimally	$\frac{1}{2} = \cdot 4$
°	„	$\frac{1}{3} = \cdot 25$	.....	„	$\frac{1}{3} = \cdot 3$
°	„	$\frac{1}{4} = \cdot 1666$	.....	„	$\frac{1}{4} = \cdot 2$
°	„	$\frac{1}{5} = \cdot 125$	.....	„	$\frac{1}{5} = \cdot 15$
°	„	$\frac{1}{6} = \cdot 111$	.....	„	$\frac{1}{6} = \cdot 14$

ing this system of notation by tens, in applying it to fractions: thus, *one remove* to the **right** of the unit's figure stands for so many *tenths*, *two removes* for so many *hundredths*, and so on.

It is to the credit of the framers of the metric system that they adopted this scale in the building up of their weights and measures, so that *one remove to the left* of the unit means *ten times* that unit, and *one to the right* means the *tenth part* of that unit, whatever it may be. Further, they invented a set of prefixes which at once indicate the number of removes from the unit, either to the right or to the left. Thus the multiples **deca**, **hecto**, **kilo** stand for the first, second, and third removes to the left from the unit, and therefore indicate 10 times, 100 times, and 1,000<sup>th</sup> times the unit respectively. The submultiples **deci**, **centi**, **milli** stand respectively for the 10th part, the 100th part, and the 1,000th part of the unit. This will be clear from the following diagram.

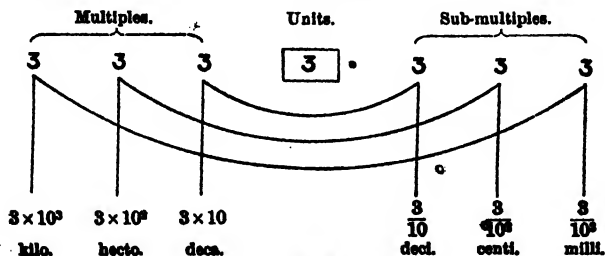


Fig. 3.

Having learned this simple rule, the connection between the metric system and the ordinary method of writing numbers is easy to understand and to remember; and one great advantage of the metric

System proceeds from this fact that it is a decimal system, the position value being indicated by the prefix: thus, *milligram* means the thousandth part of a gram, whatever the gram may be; and *kilometre* signifies a thousand metres, let the metre be what it may.

Table of Linear Measure  
(metric system):—

10 millimetres	= 1 centimetre	=	·01
10 centimetres	= 1 decimetre	=	·1
10 decimetres	= 1 metre	=	1
10 metres	= 1 decametre	=	10
10 decametres	= 1 hectometre	=	100
10 hectometres	= 1 kilometre	=	1000

#### Connection between the English and French Measures of Length.

—The metre, as will be seen from an examination of the metre scale, is a little longer than the English yard; so that if a French draper sold calico at the same price per metre as the English draper does per yard, calico would be cheaper in Paris than in London. Strictly speaking, a *mètre* is 39·3708 inches, or, less accurately, it is 3 feet 3 inches and  $\frac{3}{16}$ ths of an inch (3 feet 3·3 inches), which is the most convenient form for remembering.

In the diagram above is given a decimetre divided also into inches, from which it is apparent that the French decimetre is nearly four English inches;

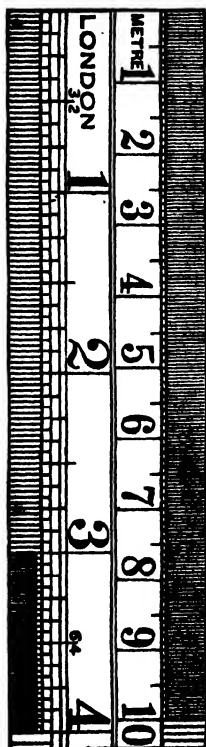


Fig. —Centimetres and Inches.

hence the metre should equal nearly forty inches, which, as we have already seen, it does.

### PRACTICAL WORK.

1. Examine the metric scale, and observe that it has 39 full inches, three-tenths of an inch, and about three-quarters of another tenth marked upon it—that is, its length in inches is 39.3708. Notice this every time you use the scale, until you are thoroughly acquainted with the relationship between the yard and the metre (3 feet 3.3 inches will be near enough and easy to remember).

2. Find the value in French measure of the yard and the inch, and in English measure of the decimetre, the centimetre, etc.

3. Measure your desk, the table, the blackboard, and the length of your classroom, and record your results in both metric and English measures.

4. Take two pieces of string or tape of unequal length, say 8.4 cm. and 6.8 cm. Measure each separately. Put the two pieces of string together, and exactly take their combined length on a third piece of string. Show that this equals 15.2 cm., which could have been got by adding the two lengths thus:

8.4 cm.

6.8 cm.

15.2 cm.

and so deduce the rule for the addition of decimals.

5. Now cut off a piece from either of the two measured pieces of string. Measure the piece cut off, and also the remainder, and show that what we have done is the same as the subtraction of the piece cut

off from the original length. Record your results thus :—

$$\begin{array}{rcl}
 \text{Original length of string} & = & 8.4 \text{ cm.} \\
 \text{Length after cutting} & = & 5.9 \text{ cm.} \\
 \hline
 \text{Length cut off} & = & 2.5 \text{ cm.}
 \end{array}$$

and so deduce the rule for the subtraction of decimals.

6. Take a piece of tape and fold it up until it is fourfold. Measure the doubled-up length, and multiply this length by four. Unfold the tape, and measure the whole length, and so deduce the rule for the multiplication of decimals by whole numbers. Record your results thus :—

$$\begin{array}{rcl}
 \text{Length of fold} & = & 4.7 \text{ cm.} \\
 \text{Number of folds} & = & 4 \\
 \hline
 \text{Length of tape} & = & 18.8 \text{ cm.}
 \end{array}
 \left. \begin{array}{l} \\ \\ \end{array} \right\} \begin{array}{l} \text{Actual measurement} \\ \text{of whole length of} \\ \text{tape} \end{array} \Bigg\} = 18.8 \text{ cm.}$$

7. Take another, or the same length of tape, and first measure its length. Now fold it up into a convenient number of folds, say four. Divide the whole length by the number of folds, and also measure the length of each fold. Show that the two results agree, and so deduce the rule for the division of decimals by whole numbers. Record your results thus :—

$$\begin{array}{rcl}
 \text{Total length} & = & 16.8 \text{ cm.} \\
 \text{Number of folds} & = & 4 \\
 \text{Length of fold} & = & 4 \overline{) 16.8} \\
 & & \bullet \quad 4.2 \text{ cm.}
 \end{array}
 \left. \begin{array}{l} \\ \\ \end{array} \right\} \begin{array}{l} \text{Actually measured length} \\ \text{of each fold} \end{array} \Bigg\} = 4.2 \text{ cm.}$$

## II. WEIGHT AND MASS.

It was the great English philosopher, Sir Isaac Newton, who first established the important fact that every particle of matter in the universe attracts every other particle of matter with a force that is greater

the less the distance between the attracting particles, and also greater the greater the amount of matter in the bodies. This attraction is of much importance, since to it is due the property of matter which we name **weight**. The huge earth attracts all bodies, and this **earth-pull is weight**. Thus the weight of a body depends not upon itself but upon the presence of other bodies, between which and itself there is a mutual attraction. Hence all bodies near the earth, being less in mass than the earth, are pulled towards it, in accordance with this law of gravitation; and it is this pull, as above stated, that constitutes the weight of the body. If a body were to be alone in space, that body would have no weight whatever; for which reason this property of weight is to be regarded as an accidental one, liable to fluctuation, and depending upon the presence of other bodies for its existence.

**Mass and Weight.**—If we examine a sphere of wood and one of lead, equal in size, it does not take long to

discover that the lead is much heavier than the wood. The converse of this is shown to be true in the accompanying diagrams, in which equal masses are shown to be very unequal in size or volume.

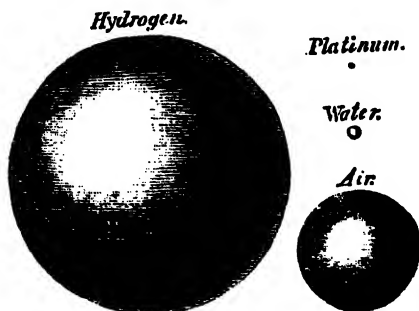


Fig. 5.—Equal mass and unequal volume.

We have already stated that weight is simply the pull of

the earth upon a body; and as the force of the pull depends upon the quantity of matter in the attracting bodies, there is evidently more matter in the ball of lead than in the wooden one of the same size. We express this by saying that the mass of the leaden ball is greater than the mass of the wooden one, for the term **mass signifies quantity of matter**. It will be seen that we cannot compare unequal masses by merely placing them side by side. We must compare them by weighing them. This renders it all the more necessary that the difference between mass and weight should be clearly apprehended. Weight, to reiterate our former statement, is merely the pull of the earth upon the body; whereas mass is the quantity of matter in the body. The former increases as we approach nearer to, and decreases as we recede further from, the centre of the earth; the latter, of course, remains constant at these varying distances, for by merely removing a body further from, or by taking it nearer to, the centre of the earth, it is clear that we do not add any matter to it, nor do we take any from it. Thus a piece of lead weighing 190 ounces at the equator would weigh an ounce more, or 191 ounces in all, at the poles, for the simple and sufficient reason that we are thirteen miles nearer the earth's centre in the latter than in the former position, and the nearer the earth the greater the pull. Although, then, we may compare masses by comparing weights, weight

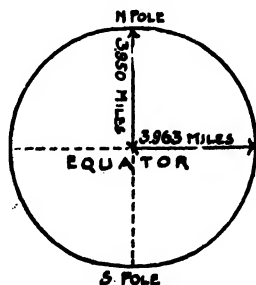


FIG. 6.



and mass are two totally different things; and to ensure the accuracy of this comparison of masses by weighing, it is necessary that the bodies should be weighed at the same place and under exactly similar conditions. Their distance from the earth's centre being the same, any difference in weight must be due to difference in mass; and so, although weight is not mass, it may, under proper conditions, be made to measure mass.

The unit of mass in the English system is the pound. It is the weight of a cylinder of platinum, 1.35 inches in height, and 1.15 inches in diameter, which is carefully preserved in the Standards Office. The multiples and sub-multiples of the pound are too well known to need mention. Concerning their selection, exactly the same remarks may be made as in respect to the linear multiples and submultiples.

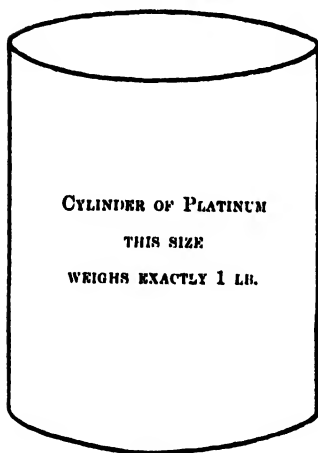


Fig. 7.

The French or metric unit of mass is the kilogram, and we can now draw attention to another advantage connected with the metric system—namely, the facility with which we can pass from one unit to another. If we take a cubic decimetre, and find out the exact weight of the water that will occupy this space, we shall have the kilogram, *for the kilogram is the weight of a cubic decimetre of distilled water at 4° C.* To

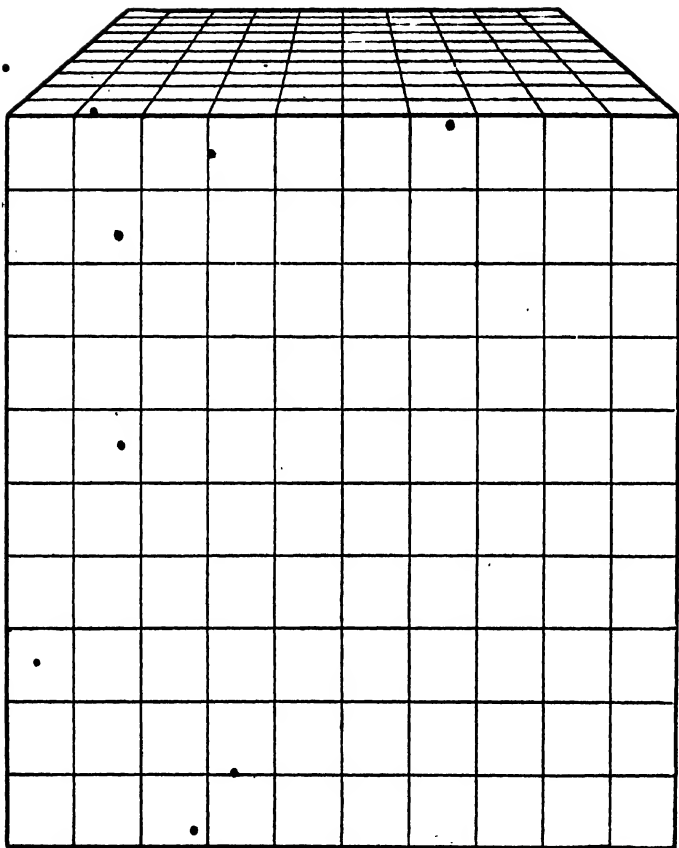


Fig. 8.—Capacity of 1 litre (1,000 cubic centimetres). Holds 1,000 grams, or 1 kilogram, of water at 4° C.

One decimetre = 10 centimetres =  $\frac{1}{10}$  of a metre = nearly 4 inches.

get the pound from the unit of length is not so easy : a cubic foot of water weighs  $62\frac{1}{2}$  lbs., and a gallon of water weighs 10 lbs.

## PRACTICAL WORK.

For this exercise there will be required—(a) metric scale; (b) darning-needle; (c) bags of shot whose weights are in the proportion 1 : 2 : 3 : 4 : 5 : 6—say 2 grams, 4 grams, 6 grams, 8 grams, 10 grams, and 12 grams respectively.

1. Pierce the metric scale at the middle, and adjust it on the darning-needle as a pivot, so that it just balances. Now use the various weights at varying distances, altering the weights and the dis-

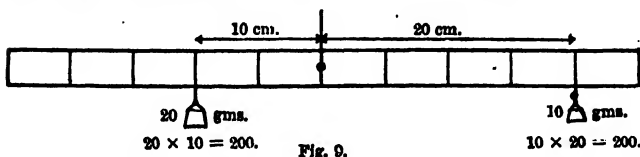


Fig. 9.

tances until equilibrium is in each case arrived at. Next multiply the distance from the centre of suspension to the place where the bag is hung by the number representing the bag in question, and record your results as follows:—

Right-Hand Side.			Left-Hand Side.		
Weight of Bag.	Distance.	Product.	Weight of Bag.	Distance.	Product.
10 grms.	10 cm.	100	20 grms.	5 cm.	100
60 grms.	5 cm.	300	10 grms.	30 cm.	300
80 grms.	4 cm.	320	16 grms.	20 cm.	320

• From this we learn that whenever there is equilibrium, the product of the weight into the distance from the point of suspension to the point of application is on each side equal. By adjusting the weights, we may, if the lever is only long enough, make any weight, however small, counterpoise any weight, however large.

2. Now make the distances equal, and add the necessary weights until equilibrium is obtained. The weights will in this case always be found to be equal.

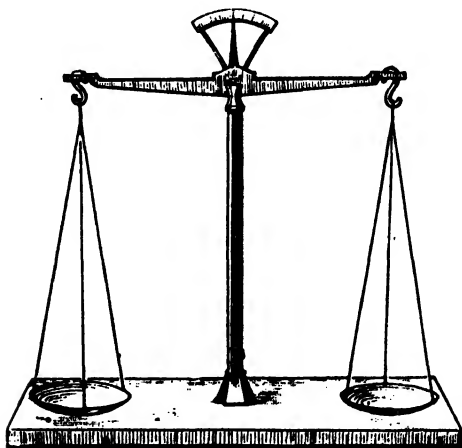


Fig. 10.—Scales or beam balance.

This is the principle of the scales or beam balance, the arms of which are carefully made exactly equal in length, so that when in equilibrium the weights on the pans must be also equal.

3. Prove this by measuring the arms of your balance.

*Rules for using the Balance.*

(a) By adjusting the levelling screws, make the stand of the balance perfectly level, as shown by the spirit-level.

(b) Put the weights on the right-hand pan, and the substance to be weighed on the left.

(c) Neither remove nor add weights until the pans are brought to a stand-still.

(d) Start with the heaviest weight, and proceed step by step downwards until the right weight has been ascertained—that is, until equilibrium has been reached.

(e) Notice what weights are absent from the box. This will tell you the weight of the body. Confirm this by adding up the weights as you replace them in the box.

(f) Remember that the balance is a delicate instrument, to be handled with extreme care.

4. Weigh same sized cubes of oak, ash, pine, poplar, etc., and record your weighings on squared paper, each square standing for a definite weight: thus:—

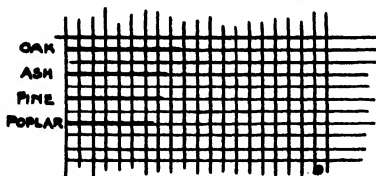


Fig. 11.

### III. AREA.

In the calico spoken of in the first lesson we considered only the length, although we could not easily get out of our minds the fact that it had breadth as

well as length; and this breadth must certainly be kept in mind, if we would form a correct idea of the real size of the piece of calico. How can we think of both length and breadth in conjunction? By thinking of how much space the calico would cover—that is, by regarding neither the length nor the breadth separately, but by thinking of the **surface-size**, in which the ideas of length and breadth merge. This surface-size, or covering power, is called **area**. Consider the following diagram.

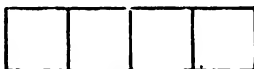


Fig. 12.

We have one tier of four squares, because the base is four times the length of each square side, and the height is equal to the side of the square. Now put on another tier, and again count the squares thus obtained.

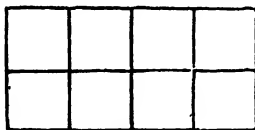


Fig. 13.

There are eight of them. If we were to put on another tier, we should have twelve of them, and so on. It is evident that, in the case of a rectangular figure such as this, the surface comprises as many squares as equal the product of the number of divisions in the length by the number of divisions in the breadth. Hence the rule for obtaining the area of a rectangular surface—namely, **multiply the length by the breadth**. At the end of this chapter are given the

rules for finding the area of various surfaces. The student should verify these by actually measuring the surfaces—for example, after having measured the base and the height of a triangle, the triangle should be marked off on sectional paper, in order to see that the squares enclosed by the triangle equal the value got by multiplying the base into one-half the perpendicular height.

**Area by Weighing.**—One of the most interesting ways of proving the truth of the various rules for ascertaining areas is by cardboard weighing, the principle of which operation is very simple. If we get cardboard which is of the same kind throughout its substance, then equal weights of it must correspond to equal areas—for example, suppose a square of this cardboard and a triangle are found each to weigh the same, we infer that the area of the square is exactly that of the triangle.

### PRACTICAL WORK.

1. Take a piece of string and double it up four or five times. Now measure carefully the length of the folded-up string; after which, unfold it and measure the entire length, and show that this equals the product of the length of the folded-up string by the number of folds. Thus, if, when folded up, the string is 8.4 cm.—that is, 84 mm.—long, its entire length should be, supposing four folds to be taken, 33.6 cm., or 336 mm. long. This gives the rule for the multiplication of decimals by whole numbers,—namely, multiply as in whole numbers, and mark off as many decimal places in the product as there are decimal places in the multiplicand.

2. Draw a line 3·6 cm.—that is, 36 mm.—long, and another one at right angles thereto 2·3 cm.—that is, 23 mm.—long. Complete the rectangle. Now

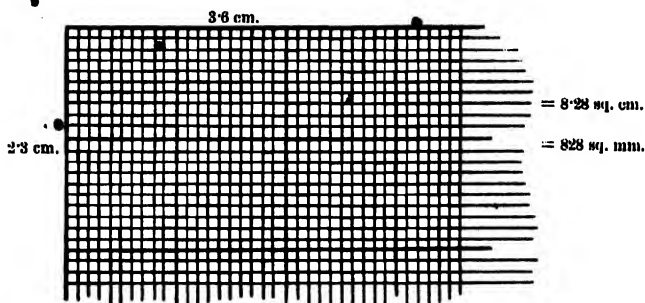


Fig. 14.

count the centimetre squares, and proceed to ascertain the area of the parts not included in these squares: thus, 6 mm.  $\times$  23 mm. = 138 sq. mm., and 30 mm.  $\times$  3 mm. = 90 sq. mm., or a total of 228 sq. mm., which, as 100 sq. mm. make a sq. cm., gives 2·28 sq. cm. + 6 sq. cm.—that is, 8·28 sq. cm. as the area of the 3·6 rectangle. But multiplying by 2·3 yields the same result; thus proving the rule for the multiplication of decimals—namely, multiply as in whole numbers, and mark off as many decimal places in the product as will equal the sum of the number of decimal places in both multiplier and multiplicand.

If the sides had been 1 mm. respectively, the area would have been 1 sq. mm.—that is, if ·1 cm. be multiplied by ·1 cm. the product is ·01 sq. cm. If the sides were ·01 cm. and ·1 cm. respectively, the area would be ·001 sq. cm.; and again the rule for the multiplication of decimals is illustrated.



3. Cut out a square of cardboard, and weigh it with great care. Now measure the base, and so get its area by squaring this base. Record your results thus:—

$$x \text{ sq. cm. of the cardboard} = y \text{ gms.}$$

4. Now cut out various figures in the same cardboard, such as parallelogram, triangle, circle, hexagon, and having most carefully weighed them, proceed to determine their area by simple proportion.

In order to test your work, the following rules for determining the area of various figures are given:—

- Area of square = square of base.
- Area of rectangle = base  $\times$  height.
- Area of triangle = base  $\times \frac{1}{2}$  height.
- Area of parallelogram = base  $\times$  perpendicular height.
- Area of circle =  $\pi r^2$   $\times$  square of radius.
- Area of hexagon = base  $\times 3$  times height of equilateral triangle upon the base.

#### IV. VOLUME.

We have now to advance one step further in our measurement of size. If we take a box or hollow cube, and wish to express the size of it, we cannot do so by saying how long it is; nor have we done this when we state the breadth also. A further dimension—namely, depth—must be known before the real size of the box can be stated.

Another way of expressing the size of a vessel such as the one in question is by stating how much liquid it is capable of holding. The size of a vessel is called its **volume**, **capacity**, or **holding power**.

**Unit of Volume.**—In England the unit of volume is

the **cubic yard**. In the metric system it is the **cubic metre**. But each of these legal units is much too large, and consequently the customary unit in England is the **cubic foot**, and in France the **cubic decimetre**. (See Fig. 8, p. 21.)

In addition to this unit of cubical capacity, there is the unit of liquid measure, which with us is the **gallon**, and in France the **litre**, the volume of a cubic decimetre. The gallon is the volume of ten pounds of water at  $4^{\circ}\text{C}$ ., so that a gallon of water weighs ten pounds. A litre of water weighs a kilogram, or a thousand grams, and therefore the gram is the weight of the thousandth part of a litre of water—that is, the weight of a cubic centimetre of water at  $4^{\circ}\text{C}$ .

To secure accurate measurement of volume, many kinds of measuring vessels have been devised, of which the chief are (1) the **graduated cylinder**, (2) the **burette**, (3) the **pipette**, and (4) the **measuring flask**.

These vessels are all figured below, but they should themselves be examined carefully, so that the student may realize what they are, and when they are to be used.

The **graduated cylinder** is divided either into cubic centimetres and its subdivisions, or into ounces. Thus a cylinder holding half a litre will be divided into 500 c.c., each of which will be further subdivided into tenths. A cylinder holding a pint, which is the eighth part of a gallon, will be divided into 20 oz., each again being subdivided into tenths. As the gallon of water weighs 10 pounds, or 160 ounces, it will be



Fig. 15.  
Graduated  
cylinder.

seen that the fluid ounce, as it is called, will be the volume taken up by the twentieth part of a pint; and it is further to be noted that the only fluid ounce that really weighs an ounce, is an ounce of water, so that the term fluid ounce is somewhat misleading, it being an ounce, or more or less than an ounce, according as the liquid is water, or a liquid heavier or lighter than water respectively.

The burette is a vessel graduated for delivering a measured portion of liquid. It is provided with a stop-cock at the bottom, and is graduated in cubic centimetres and cubic millimetres.

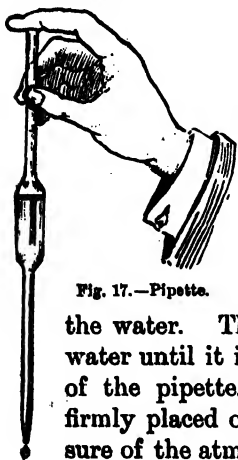


Fig. 17.—Pipette.

The pipette is a very interesting bit of apparatus for delivering certain volumes of liquid, the amount of which is stated on the pipette. Suppose it is required to deliver 50 c.c. of water: a 50 c.c. pipette is selected, and the narrow end is placed in

the water. The next thing is to suck up the water until it is past the mark in the upper part of the pipette. Thereupon the index finger is firmly placed over the aperture, and so the pressure of the atmosphere is prevented from forcing down the liquid which remains in the pipette.

Very carefully now the finger is slightly loosened, and the air being thus allowed to enter depresses the liquid. As soon as the liquid is on a level with the mark on



Burette.

the pipette, the finger is again pressed tightly down on the aperture, and the 50 c.c. of water is now entirely under control for delivery. When the finger is removed from the aperture, 50 c.c. of water will flow from the pipette into whatever vessel is placed to receive it. Incidentally it may be noted how well the pipette establishes the fact that air exerts pressure.

The graduated or measuring flask has a mark on its neck, and when full up to the mark it contains the volume recorded on the flask.

To graduate a flask in cubic centimetres, we first carefully counterpoise it, and then pour in one gram of water at  $4^{\circ}\text{C}$ . for every cubic centimetre the flask is intended to hold. Thus, if we want a half-litre flask, we take heed to have just 500 gms. of water, the level of which is marked on the neck of the flask by a file.

It is important to notice that whereas the cylinder and the flask register the amount of liquid they *contain*, the pipette and the burette register the amount of liquid they are intended to *deliver*.

**Volume by Displacement of Water.**—If a vessel is quite full of water, and a pebble is placed in it, some of the water will overflow, because where the pebble is there cannot be water, and the water that is thus displaced by the pebble must overflow, as the vessel was already full when the pebble was placed therein. If we measured the water displaced, we should know exactly the size or volume of the pebble. A better way of performing this experiment is to have a cylin-



Fig. 18.  
Measuring flask.

der or a burette only partly filled with water, so that when the solid, whose volume is required, is placed in the vessel there is no overflow, but a rise in level. The difference between the old level and the new

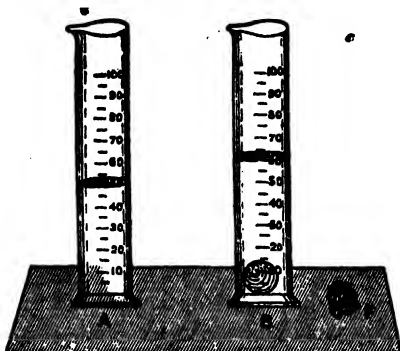


Fig. 19.—Volume by displacement of water.

exactly measures the volume of the pebble. By this simple means the volume of a most irregular solid can be found, provided, of course, the solid is not soluble in the liquid used.

### PRACTICAL WORK.

For this exercise there are required portions of soap, cheese, clay, butter, pebble, slate-pencil, and various other solids, measuring vessels, and balance.

1. Cut out a cubic centimetre of as many substances as possible—for example, soap, cheese, clay, butter—and find the exact weight in each case.

2. Find the volume, by displacement of water, of as many insoluble solids as possible—for example, pebble, piece of slate-pencil, nail, etc.—using the burette and the cylinder for this purpose.

3. Take a small Würtz flask, or any other vessel provided with a suitable exit. Fill exactly up to the exit with water, and then place a stick of slate-pencil or other solid in the water, carefully collecting and measuring the water which flows out by way of the side tube when the solid is placed in the flask.



Fig. 20.

4. Into a small, carefully - counterpoised flask deliver successively 10, 20, 30, 40, 50, etc., c.c. of water, and weigh so as to determine the weight of a c.c. of water, which, as above stated, should be about a gram.

5. Carefully weigh specimens of chalk, iron, brass, etc., the specific gravity of which is known. After weighing, proceed as follows :—

*Chalk.*

Weight of chalk taken = 8.5 gms.

If this were water, its volume would be = 8.5 c.c.

But chalk is 2.6 times as heavy as water,

∴ specific gravity of chalk taken =  $\frac{8.5}{2.6}$  c.c.

6. Obtain, by weighing, the volume of as many substances as possible, using a table of specific gravities to tell how many times the body is heavier than water. Thus, if the specific gravity or relative density of iron is 7.8, and we weigh out 50 gms. of it, we shall reason thus :—

50 gms. water measure 50 c.c.

But iron is 7.8 times as heavy as water,

∴ 50 gms. iron measure  $\frac{50}{7.8}$  c.c.

## V. DENSITY OR SPECIFIC GRAVITY.

We have already shown that although mass or quantity of matter in a body is quite different from

weight or earth-pull upon the body, still weighing is the only easy way in which we can conveniently compare masses; and if we are careful to weigh the bodies to be compared under exactly the same conditions, the comparison is a just one. Thus, if a piece of butter weighs just as much as a pound weight, we are justified in believing that there is just as much matter in the butter as there is in the weight; likewise, if a cube of lead weighs half as much as the same-sized cube of platinum, we infer that there is twice as much matter in the platinum cube as there is in the leaden, and we say that the density of platinum is twice that of lead, because, we argue, the matter is packed twice as closely in the former as in the latter case, just as we speak of a dense fog or a dense crowd, because of the closeness of the packing. To simplify matters, it has been agreed upon to always express the *density* of a body by reference to the unit of volume, which in the English system is, for scientific purposes, the cubic foot. Hence, we define **density** as the **mass of unit volume**; and as mass is expressed in terms of weight, we arrive at the simple rule that the **weight in pounds** of a cubic foot of a body represents its density, thus:—

Name of Body.	Weight of a Cubic Foot.	Density.
Water .....	62½ lbs.	62·5
Mercury .....	812 lbs.	812
Iron .....	432 lbs.	432
Oak .....	73 lbs.	73

In the metric system the unit of volume, for scientific purposes, is the cubic centimetre; wherefore the density of water in this system is 1, as a cubic centimetre of water weighs one gram.

**Specific Gravity** is another and the more usual way of comparing masses and densities. In the case of density determinations, we simply find the mass of unit volume as expressed in pounds, so that the ratio is between the pound weight as unit and the number of pounds of matter in unit volume of the substance.

Here we may take "specific" to mean "special," and "gravity" to stand for "mass expressed in weight," and so specific gravity determinations express relative mass. The speciality about the process is that we compare, not with the pound weight, but with the mass (or weight) of water equal in volume to the body compared. We first determine the weight of a certain volume of the substance under examination, and then calculate the weight of the same volume of water, and divide the former by the latter. Thus:—

$$\text{Specific gravity} = \frac{\text{Weight of substance.}}{\text{Weight of an equal volume of water.}}$$

### PRACTICAL WORK.

1. Counterpoise a light glass-stoppered bottle or flask holding about 50 c.c. Now fill up to the top with water, and insert the stopper, wiping off the water that is forced out by the insertion of the stopper. Weigh the water, and so calculate the volume of the bottle, allowing 1 c.c. for each gram of water.

2. In the same bottle weigh milk, vinegar, oil, paraffin, tea, coffee, salt water, methylated spirit, etc. Divide each of the weights thus obtained by the weight of the water as obtained in the last experiment; the quotient is the specific gravity of the



liquid in question, and expresses its density as compared with water as unity.

3. Find the weight of 1 c.c. of each of the liquids by dividing the weight of the volume required to fill the bottle by the number of cubic centimetres contained in the bottle. As a cubic centimetre of water weighs 1 gram, these numbers will be the specific gravity of the liquids in question.

This is the usual way of ascertaining the specific gravity of liquids, except that the trouble of weighing

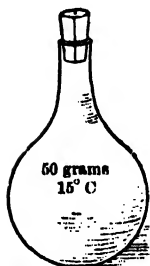


Fig. 21.—Specific gravity flask.

the water each time is avoided by using a specific gravity flask, which is specially made to hold just 50 grams of water about 15° C. The special device consists in a stopper which has a fine perforation through it, so that when it is pressed down upon the liquid, any excess escapes by way of this perforation. Instead, therefore, of weighing the water and then the liquid, we simply counterpoise the specific gravity

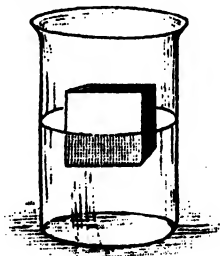
flask, then fill it up with the liquid under determination, the exact volume required being obtained by the insertion of the stopper above described. All that now requires to be done is to weigh the liquid and then divide by 50, which is the weight of the volume of water capable of filling the specific gravity flask. Thus, suppose the bottle contains 75 grams of a liquid, the specific gravity will be  $\frac{75}{50} = 1.5$ .

If possible, the specific gravity flask should be used in the above experiments in place of the ordinary flask.

## VI. SPECIFIC GRAVITY OF LIQUIDS—SECOND METHOD.

There is an even more speedy way of obtaining the specific gravity or relative density of liquids, which will be readily understood after a brief discussion of **Floating Bodies**. When a solid is completely immersed in a liquid, it displaces its own volume of the liquid, as was seen in the overflow or rise of the water in the experiments on determining volume by displacement of water.

But in the case of a ship, or a piece of wood, or other floating body, the case is different, for it is evident that the floating body does not displace its own bulk of the liquid in which it floats. How much does it displace? To answer this question we must first ascertain the weight of the floating body, for it **displaces its own weight of the liquid in which it floats**. If we float a regular solid in water so that we can readily obtain the volume of the immersed portion, it will be a very easy matter to obtain the weight of the floating body without actually weighing it. Thus, suppose a block of wood, as in Fig. 22, to float in water, so that the immersed portion is indicated by the



whole of the block, since it displaces its own weight of water, and 288 grams is the weight of water actually displaced.

**Hydrometers.**—This illustrates the principle of the hydrometer, an instrument for determining the specific gravity of liquids.

The common hydrometer consists of a long stem and a bulb of glass. In the bulb is placed some mercury, so that the instrument may float upright in the water. By experiment it is ascertained how far down the neck sinks in water, and a scratch is made on the narrow stem or neck at this point. If, when placed in liquid other than water, the hydrometer sinks below this mark, the liquid in question is *lighter than water*. If, contrariwise, the scratch mark is above the surface of the liquid, we have to do with a liquid *heavier than water*.



Fig. 23.  
Simplest  
form of  
hydrom-  
eter.

A little thought will make it easily understood why this is so, especially if what has been said in the previous section is remembered. Hydrometers are graduated to tell not only that a certain liquid is lighter or heavier than water, but also to tell exactly its specific gravity by appropriate figures upon the stem.

**Lactometer.**—This is the instrument used to determine whether or not milk has been adulterated (1) by the removal of the cream, (2) by the addition of water. The specific gravity of milk is 1.032, and a hydrometer is so constructed as to show whether the specimen under examination has this density or not. Thus, in the lactometer figured below, the mark M indicates the depth to which the instrument would sink

in a pure specimen of milk, whilst W shows the depth to which it would sink in water. This, however, is not sufficient; for, by removing the cream and adding water, the required density—namely, 1.032—may still

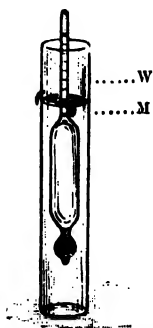


Fig. 24.—Lactometer.

be maintained, since Cream, which consists of minute globules of fat encased in albumen, is lighter than milk, and its removal would raise the specific gravity of the liquid, which could again be lowered to 1.032 by the addition of water. It is necessary, then, to ascertain the amount of cream that should be present. If the milk is placed in a narrow vessel, and is allowed to settle,

the light cream will rise to the top, and from 10 to 12 per cent. of the height of the column should be cream. A vessel is graduated in such a way as to indicate the height of the cream layer, and so the proportion of cream is calculated. If the cream has been removed, and the specific gravity is still normal, it is certain that water has been added; and even if the cream remains, but the specific gravity is below 1.032, it is again evident that water has been added. To analyze milk thoroughly, requires both chemical and microscopical examination in addition to cream and specific gravity determinations; but for ordinary household purposes the two latter determinations would suffice.



Fig. 25.—Vessel used to ascertain the amount of cream in milk.

**Nicholson's Hydrometer.**—This is a brass hydrometer with a plate or scale-pan at the top of the stem for

carrying weights, and a weighted part at the bottom for keeping the hydrometer upright as it floats. A mark is made on the stem, and when the hydrometer is floated in a liquid, sufficient weights are placed in the scale-pan to sink the mark to the level of the liquid. Thus a fixed volume of the liquid corresponding to the volume of the hydrometer up to the mark is in each experiment displaced. Hence we have a ready means of comparing the weights of equal volumes of different liquids. For example, suppose we wish to determine the relative density of olive oil

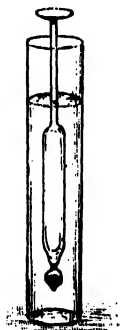


Fig. 28.—Nicholson's hydrometer.

by means of Nicholson's hydrometer, we shall (1) float the hydrometer in water, and add weights until the mark is level with the water, when the weight of the hydrometer plus the weights added are together equal to the weight of the volume of water displaced by the hydrometer. (2) We next float the hydrometer in the oil, adding weights as before until the mark is level with the surface of the oil, when the weight of the hydrometer plus the weights added to bring the mark level with the surface of the oil are together equal to the weight of the oil displaced by the hydrometer. But this volume is exactly equal to the volume of water displaced in the former experiment, and the relative density or specific gravity of the oil

$$= \frac{\text{Weight of oil.}}{\text{Weight of same volume of water.}}$$

$$\text{That is,} = \frac{\text{Weight of hydrometer + added weights in (1).}}{\text{Weight of hydrometer + added weights in (2).}}$$

## PRACTICAL WORK.

1. Take a test-tube, and pour mercury into it as ballast until it floats upright in water. Mark the tube to show how deeply it sinks in the water, and compare this with the depth to which it sinks in alcohol, turpentine, milk, vinegar, oil, and other liquids.

2. By means of the specific gravity flask, and also by the aid of the hydrometer, determine and record the specific gravity of methylated spirit, turpentine, beer, vinegar, olive oil, petroleum, and milk.

3. Put a new-laid egg in water, and then in water that has been made very salt. Notice that it floats in the salt water. What do you learn from this?

4. With hydrometer and with specific gravity flask, show that the salt water in which the egg floats is heavier than the water in which it sinks.

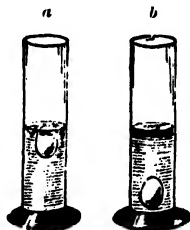


Fig. 27.—Egg in (a) salt and in (b) fresh water.

Thus the buoyancy of water depends upon its density, as is shown by the fact that a man can hardly sink in the very salt water of the Dead Sea.

5. Ascertain the depth to which the lactometer sinks in new milk, skim milk, and water respectively; also determine the amount of cream in a few specimens of skim and unskimmed milk.

6. Using Nicholson's hydrometer, determine the relative density of olive oil, as explained in the opening remarks of this chapter. Determine other liquids by the same means.

## VII. GENERAL EFFECTS OF HEAT ON MATTER.

**Heat is Power or Energy.**—At one time it was believed that heat was a material substance called **caloric**, and that "the difference between the states of a body when hot and when cold was due to the presence of a substance called caloric, which existed in greater abundance in the body when hot than when cold." But a piece of iron when heated weighs exactly the same as when cold, so that it certainly gains nothing in substance in passing from the cold to the heated condition. Heat, then, is not matter, but power or energy, which can accomplish much, as a cursory look round will show.

**Change of State.**—If a piece of solid water, or ice, as it is called, is heated, it is converted into liquid water, which again, on being sufficiently heated, becomes **water-gas**. This change of state from solid to liquid, and from liquid to gas, is one of the most common and most interesting of the many changes brought about by the power called heat.

**Expansion of Bodies.**—Heat causes bodies, whether solid, liquid, or gaseous, to expand. The cracking of the lamp glass is due to the fact that glass is a bad conductor of heat, so that when heat is applied to one part of the glass it does not readily pass to the other parts. Hence the heated portion in expanding forces itself away from the remainder, and a rupture ensues. For the same reason the ends of consecutive rails in a railway are always a short distance apart, so that when the rails are heated in summer, or by the friction of carriages at any time, they have room for expansion, and are not bent, as

would inevitably be the case were it not for this safeguarding gap.

The expansion of liquids and gases will be apparent from the experiments which follow.

The formation of rain, dew, mist, &c., and of winds and ocean currents, etc., depends upon the expansive effects of heat. In the home, ventilation is made easy and simple by the action of heat in expanding the air, and so producing currents that result in an out-flow of heated air and an influx of cooler air. It is because of this expansive effect of heat that we are careful to weigh the cubic centimetre of water at 4° C. when we want the water to weigh a gram, and why we keep the standard yard at 62° F.

**Heat causes Decomposition, and also brings about Combination.** — If some chlorate of potash is heated, oxygen comes off in consequence of the decomposition of the substance under the influence of heat; whilst the burning of fuel and luminants, such as coal-gas and petroleum, is due to the fact that the application of heat causes them to combine with the oxygen of the air. This will be dealt with fully in a later section.

**Cooking.**—By means of heat, food is made both more palatable and more digestible, as will appear later.

**Incandescence.**—When the temperature of a body is raised, it may (1) decompose, (2) combine with some other body, or (3) become white hot or incandescent. This last possibility is a most important one, since it shows that heat and light are different only in degree and not in kind; for the passage from heat to light is simply due to an increase in tem-



perature, light being emitted from bodies that have been raised to a high temperature. This is the case with the sun, with the electric light, with gas, petroleum, candles, the limelight, and indeed with all luminous bodies, which are simply solids, liquids, or gases raised to incandescence by great heat. The so-called mantle of the incandescent burner, that has lately become so popular, consists of a substance that does not easily decompose, and that therefore can be raised to the state of incandescence, to which is due the brilliant white light for which the mantle is valued.

Finally, life itself is dependent upon light and heat, so that it is impossible to overrate the importance of this great natural power that we call heat.

### PRACTICAL WORK.

1. Heat some chlorate of potash in a test tube, and place a smouldering chip at the mouth of the tube. The rekindling of the chip that results is due to the action of the oxygen produced by the decomposition of the chlorate of potash.

2. Take a piece of iron wire, and heat it in the flame of a Bunsen burner or of a spirit-lamp. Observe the change from black heat to red heat, and finally to white heat. If the wire be carefully examined after heating, it will be found that a scale has formed on it during the heating process. This is a new body, due to the joining together of the wire and the oxygen of the air.

3. Next heat a piece of platinum wire, and note the same changes; and also that the platinum wire upon cooling is found to be totally unchanged, the

heat having brought about no union between the platinum and the oxygen of the air.

4. Take now some magnesium wire, and heat this. Observe the brilliant light and the white product of the burning. How totally unlike the magnesium it really is! It is an entirely new body—namely, magnesium-oxide. Here the heat has brought about complete combination between the metal and the oxygen of the air.

5. **Expansion of Solids.**—Take a flat piece of iron about a foot and a half long. Place it upon two wooden blocks, as in the figure, one end being kept firmly in place by a heavy weight, the other resting upon a needle placed upon the smooth surface of

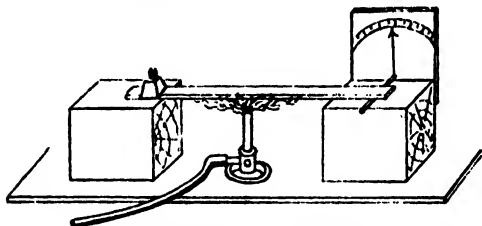


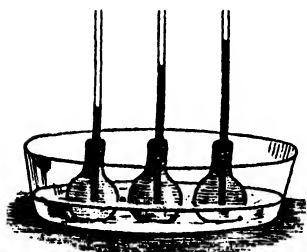
Fig. 28.—Expansion of a metal bar.

the block, the needle having a straw attached to it by means of wax. A divided semicircle of cardboard is placed near the block, so that it may measure any movement of the needle. Let the straw be vertical at the commencement of the experiment. Now proceed to heat the iron, and observe the movement of the straw, and therefore of the needle, as the iron becomes hot. This movement is due to the expansion of the iron; and the object of the straw is to make this movement more apparent than would be the

case if we merely noticed the roll of the needle, or the lengthening of the bar, which is very small indeed.

Strips of other metals than iron may be used, and similar results will be obtained.

**6. Expansion of Liquids.**—Take three 4-oz. flasks. In the cork with which each of the flasks is to be fitted insert about a foot of glass tubing of  $\frac{1}{4}$ -inch bore. Fill the flasks with water, turpentine, and alcohol respectively, and push in the corks until the liquid stands at exactly the same height in each flask. Place the three flasks in a vessel containing



*Water, Turpentine, Alcohol.*  
Fig. 29. —Expansion of liquids.

water, say at about  $80^{\circ}\text{C}$ .

At first the liquids sink slightly. This is due to the expansion of the glass before the heat has had time to expand the liquids. But soon the liquids are heated and expand, as is seen from the columns that begin to rise in the glass

tubes. The alcohol is evidently most expansive; and next to that the turpentine, the level of which rises nearly twice as high as the level of the water. From this experiment we learn, not only that liquids expand, but also that for the same increase in temperature there is not the same amount of expansion for each liquid, but a different amount of expansion for each different liquid. Volatile liquids, such as alcohol and chloroform, are, as a rule, most expansive. This is well seen by a comparison between the expansions of quicksilver and alcohol; the former

liquid will be seen to expand six times less than the latter.

7. **Expansion of Gases.**—Fit a good-sized tin can with a leading tube, the end of which dips under a beehive shelf in a trough filled with water, as in the diagram. Over the beehive shelf place a gas jar filled with water, and apply heat to the can. Observe that bubbles of gas soon pass up the gas jar and drive down the water. These are air bubbles consequent upon the expansion of the air in the can by the heat, which causes some of the air to be driven out.

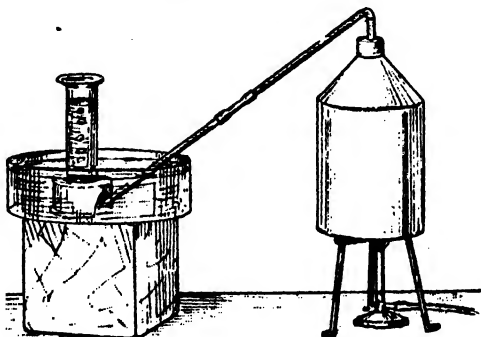


Fig. 30.—Expansion of gases.

By-and-by, if the heat is reduced, the water will rush up and fill the space previously filled by the air now collected in the jar, affording a proof of the contraction of gases upon cooling.

8. Fit a cork, through which passes a straight tube, into the neck of a 2-oz. flask. Invert the flask, and pass the tube through the cork in the neck of a wide-mouthed bottle containing water coloured with litmus or other colouring matter. Warm the flask

with the hand to expel some of the air. This air will bubble up through the coloured water. Now remove the hand, and allow the temperature of the flask



Fig 31.  
Air thermometer.

to sink. Observe the column of coloured water that rises in the tube. If the upper flask is further cooled, this additional loss of heat will be registered by a further increase in the height of the column of coloured water. If, on the other hand, a further expansion of the air in the flask is brought about by the application of more heat, then the column of coloured liquid will be observed to fall. This arrangement serves the purpose of a temperature measurer or thermometer; and as the column whose contraction and expansion causes the rise and fall of the coloured water is an air column, this arrangement of flasks is an **air thermometer**.

It is worthy of note that, whereas solids and liquids expand differently for the same increase of temperature, the increase of volume for the same increase of temperature is practically the same for all gases, which expand  $\frac{1}{273}$  of their volume at  $0^{\circ} \text{C}$ . for each additional increase of  $1^{\circ} \text{C}$ . in the temperature: thus, a litre of air at  $0^{\circ} \text{C}$ . would become two litres at  $273^{\circ} \text{C}$ .

## VIII. THE THERMOMETER.

**Construction of Thermometer.**—The air thermometer constructed in our last experiment would be of little

use in determining either the temperature of our rooms or of boiling water. For such purposes we make use of a mercury thermometer, or, if the temperature to be measured is not a high one, an alcohol thermometer. In the construction of a thermometer we have (1) to fill the tube; (2) to fix the boiling point and the freezing point of water; and (3) to graduate the interval between these two important thermometric points. To construct a mercury thermometer is too difficult for the beginner to attempt; but all the main points may be learned just as well from an alcohol thermometer, the construction of which is comparatively simple.

### PRACTICAL WORK.

A piece of thermometer tubing of  $\frac{1}{32}$ -inch bore is to be taken, and a bulb blown on to the end. Now connect the end remote from the bulb with a funnel by means of indiarubber tubing, and place in the funnel some methylated spirit, coloured with aniline. Surround the bulb with hot water by placing it in a vessel containing water that has been heated. The air bubbles out and escapes through the alcohol, in consequence of the expansion due to the heating. The bulb is next to be removed from the water and allowed to cool, whereupon some of the alcohol will be forced into the bulb. Again insert the bulb in the hot water, when the alcohol will vaporize and drive out the air. See that the alcohol fills the bulb and that no bubbles of air remain in it. If



Fig. 32.—Thermometer tubes.

necessary, repeat the whole operation until this condition has been assured. Place the bulb in your hand, in your mouth, in warm milk, etc., and record the temperatures thus obtained. As alcohol boils at  $78^{\circ}\text{C}$ ., the thermometer thus improvised will only be available for temperatures lower than this.

**Application to Mercurial Thermometers.**—With a mercurial thermometer, we should next proceed to seal up the open end whilst mercury and mercury vapour fill its whole length, so that the sealed tube may contain



Fig. 33.  
Introducing mercury into  
thermometer tube.



Fig. 34.  
Mode of fixing the  
freezing point.

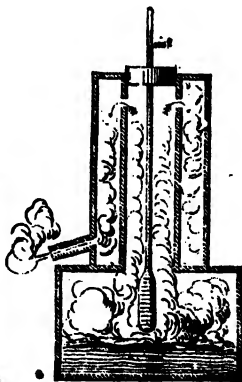


Fig. 35.  
Mode of fixing the  
boiling point.

no air, but simply mercury and mercury vapour. To find the freezing point of water, the bulb is placed in melting ice; whilst for the boiling point, the bulb is placed in steam. The accompanying illustration will show how this is to be done.

• **Graduation.**—Having found the freezing point and the boiling point of water, it now remains to divide the distance between these two points in such a way as to indicate the various temperatures between these two points, and also above and below them.

There are two chief ways in which this is done.

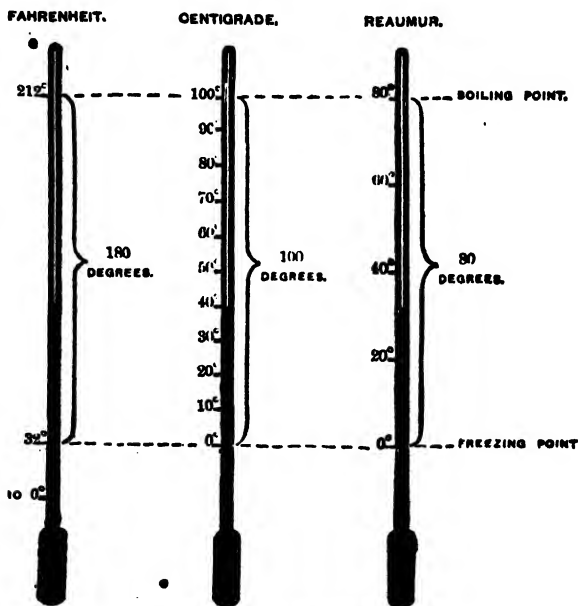


Fig. 35.—Thermometer scales.

In this country we divide the interval between the freezing point and boiling point of water into 180 small divisions or degrees. On this scale the freezing point of water is marked 32°, and so the boiling point will be 212°. The reason for this is that Fahrenheit, to whom is due this mode of graduation,



erroneously thought that  $32^{\circ}$  below the freezing point of water marked the lowest temperature attainable, which point was naturally termed zero; and consequently the freezing point of water, being  $32^{\circ}$  above this zero, was marked  $32^{\circ}$ , and adding this 32 to 180 we arrive at  $212^{\circ}$  F. as the boiling point of water, as above said.

A much more convenient mode of procedure is due to Celsius. Here  $0^{\circ}$  marks the freezing point and  $100^{\circ}$  the boiling point of water. Thus there are 100 steps or grades from the freezing to the boiling point of water, on which account it is usually known as the Centigrade thermometer. It is used exclusively in France and in other countries. In all countries scientists make use of it on account of the convenient division into 100 steps.

In reducing one scale reading to that of another, we have—

$$\begin{aligned} 100^{\circ} \text{ C.} &= 180^{\circ} \text{ F.} \\ \therefore 5^{\circ} \text{ C.} &= 9^{\circ} \text{ F.} \end{aligned}$$

As the freezing point of water is  $0^{\circ}$  C. and  $32^{\circ}$  F., we must add  $32^{\circ}$  to our Centigrade degrees when we have reduced them to Fahrenheit degrees; and we must subtract  $32^{\circ}$  from our Fahrenheit degrees before turning them into Centigrade degrees, thus:—

$$\text{C.} = \frac{5}{9} (\text{F.} - 32)$$

$$\text{F.} - 32 = \frac{9}{5} \text{ C.}$$

Fig. 36 will make clear why 32 has to be added to or subtracted from the Fahrenheit degrees.

**The Clinical Thermometer.**—This is the name given to the thermometer used by the doctor at the bedside of the patient (Gr. *κλινη*, couch). To understand the

use of this special form of temperature measure, it is only necessary to remember that **the temperature of the healthy body is always  $98.5^{\circ}$  F.** Any degree of heat above or below this is an indication of indisposition, hence the importance attached to temperature by the medical practitioner. For the rest, the clinical thermometer is to be exceedingly sensitive; hence the bulb is of the thinnest glass, and as it is required to register temperatures a little above or a little below the normal  $98.5^{\circ}$  F., its range is a short one—namely, from about  $80^{\circ}$  F. to  $120^{\circ}$  F.



Fig. 37.  
Clinical  
thermo-  
meter.

**In the house** the thermometer is not used nearly so much as it should be. In the first place, no living room should be without its thermometer. How often do we find a huge fire built up, and a dangerously high temperature thus brought about, to be succeeded, when the fire goes down, by an equally dangerous low temperature; whereas it should be the aim of every wise housewife to maintain an equable temperature of about  $60^{\circ}$  F., and to do this a thermometer is absolutely essential. Its use would certainly result in a saving of both coal and health. Again, in cooking operations, the thermometer should always be at hand, so that the temperature of oven and saucepan may be known. No good cook can afford to dispense with the aid of the thermometer.

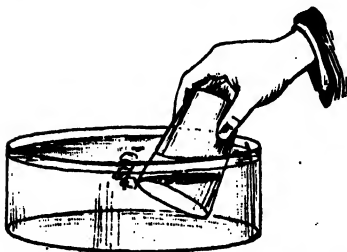
### PRACTICAL WORK.

1. With a thermometer ascertain the temperature (a) of the room; (b) of the oven; (c) of the water in a saucepan, wherein are being boiled meat, potatoes, etc.

2. Place a clinical thermometer, or, failing this, an ordinary thermometer taken from its stand, in the mouth or under the arm-pit. Note the temperature recorded.

## IX. WEIGHT OF THE ATMOSPHERE.

When we speak of a jug or a glass as "empty," we are not speaking in strict accordance with fact; for



it is only necessary to invert the vessel, and to press it down into water, in order to prove that there is something present that prevents the water from filling up the glass or the bottle.

Fig. 33.—Presence of air in an "empty" glass.

Further, if the vessel is tilted a little on one side, bubbles will flow upwards through the water and escape from the surface. Thus an "empty" vessel is really full of air; and to get a really empty vessel is a matter of no small difficulty, for even with the air-pump only a partial vacuum can be obtained. Air, being matter, is attracted by and attracts the earth—that is, it has weight. To prove this is the object of our next experiments.

### PRACTICAL WORK.

1. Take an ordinary round-bottomed Florence flask, and place therein a small quantity of water—50 c.c. will be a suitable quantity. Select a well-fitting cork, and boil the water in the flask, and so drive out all

the air, the place of which will be taken by the water vapour. Whilst the water is still boiling, insert the cork in the flask, and carefully weigh. Now remove the cork, and, as the flask cools down, the air will enter. Weigh again with utmost care, and ascertain the increase in weight, which will be the weight of the volume of air required to fill the flask. Next, by means of the graduated cylinder or the burette, find out the volume of the flask, making allowance for the water still remaining therein, and therefore of the air whose weight has just been ascertained. In this way the weight of a cubic centimetre and of a cubic foot of air should be calculated. Note that a cubic foot of air weighs about an ounce. Better results will be obtained by substituting a cork fitted with a glass tube, to which is attached a piece of rubber tubing carrying a well-fitting clip, which remains open whilst the water is boiling, but is closed immediately it is adjudged that all the air has been expelled. (See Fig. 39.)

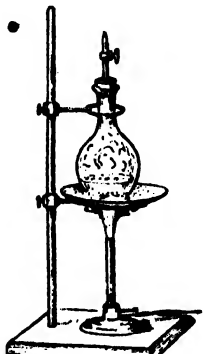


Fig. 39. - Expelling air from water.

2. Suspend a flask from one of the arms of a balance, and exactly counterpoise it with small shot or with tinfoil. Now with a Bunsen burner or a spirit-lamp heat the flask, and note that the counterpoise is now heavier than the air that fills the flask.

From this it is evident that hot air is lighter than cold air, and that gases expand on heating.

## **X. VENTILATION AND CIRCULATION IN HOT-WATER SYSTEMS.**

To the housewife, by far the most important application of the expansion of bodies on heating is in connection with **ventilation** and the **circulation of heated water**, about each of which a few words must be said.

**Ventilation.**—The sad story of the Black Hole of Calcutta always comes to mind when the dire effects of bad, or no, ventilation are under discussion. The one hundred and twenty-three victims of Suraj-ad-Dowlah's savagery furnish us with an awful illustration of the harm that may be done by impure air. How air becomes impure will be fully discussed in Stage II. Here we need only say that every living creature breathes in life-exciting oxygen, and breathes out suffocating carbon dioxide. Although a certain small proportion of this last-named gas may be present in the atmosphere without the production of any ill result, deadly indeed are the effects consequent upon an increase in the proportion of this suffocating gas in the air that we breathe. Hence the vital necessity for constantly changing the air of dwelling and school rooms, a process which is comparatively easy, thanks to the expansion of gases upon heating.

Expired or breathed-out air is at the temperature of the body—namely,  $98.5^{\circ}$  F., and therefore much warmer than the air of our rooms, which should be about  $60^{\circ}$  F. This foul air will, upon leaving our lungs, rise to the top of the room; wherefore there should be, in any intelligent scheme of ventilation, (1) **an outlet near the top of the room.** To supply the place

of this escaping impure air, there must be (2) an inlet near the floor for the admission of pure air.

The window and the chimney are exceedingly useful here. By opening the window top and bottom both (1) and (2) are secured. The inlet may become a very uncomfortable draught if proper precautions are not taken to prevent this. If the

space at the bottom of the open window be closed by a piece of wood, the space between the two sashes becomes the inlet, as shown in the diagram, where by a draught is prevented. The wise housewife will certainly not omit to make full use of this exceedingly simple and effective device for securing a proper supply of pure air. The chimney, again, is an excellent exit for

foul air when the fire is alight; and it may afford entrance to fresh air when no fire is there, hence a chimney should never be closed.

Frequently an outlet is provided at the top of the room by means of an entrance into the chimney; but there must be a valve provided to prevent the ingress of soot and smoke. An inlet is also sometimes made in

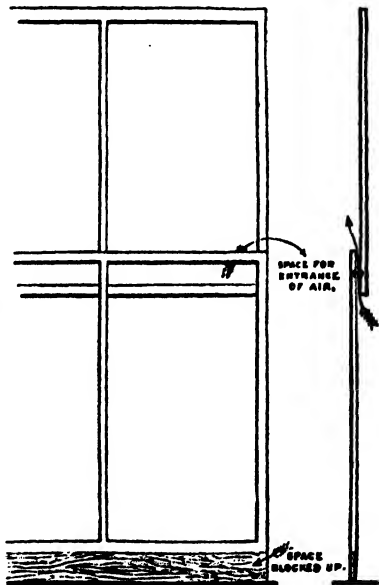


Fig. 40.—A simple mode of ventilation.

the fireplace, so that air from without may pass through the fire, and thus be warmed before passing into the room.

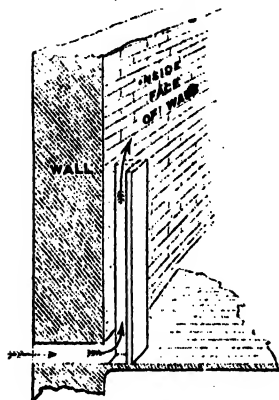


Fig. 41.  
Tobin's tube ventilator.

Another draught-preventing device is due to Tobin, whose tubes (for which see Fig. 41) are much used. For schools, Kite has pipes connected with the outside, and opening into the inside by means of warmed inlet shafts. To provide outlet, a special louvre arrangement is let into the roof.

By this arrangement the room has a complete change of air three to four times

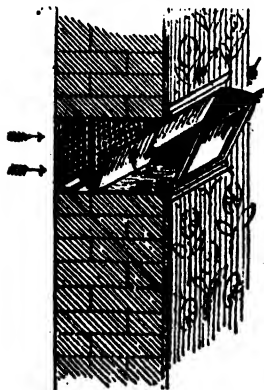


Fig. 42.—Kite's inlet and outlet.

per hour; for when the outside temperature is lower than the inside, which is normally the case,

there is a steady flow from without to within by way of the pipes and the shaft, and from within to without by way of the roof outlet.

### PRACTICAL WORK. •

1. With the flame of a lighted candle show that there is an incoming current—as shown by the inward deflection of the flame—at the keyhole and at the bottom of doors and windows, and an outgoing current at the tops of windows and doors and by way of the chimney.

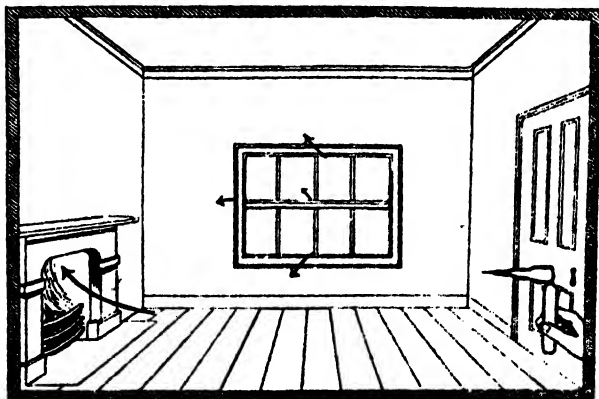


Fig. 43.—Incoming currents of air at keyhole and chinks of window.  
Outgoing current by way of chimney.

2. Mount to the top of a pair of steps and examine the air near the ceiling of a room wherein gas or lamps are burning. Notice how hot it is, and how close it seems. This shows how necessary it is that the outlet should be at the top.

3. Show models of a Tobin tube, of a louvre, and a window, etc., and explain their working.



**Circulation of Hot Water.**—In most of our houses, at least in those of the better class, baths provided with hot and cold water are to be found ; and it is all-important

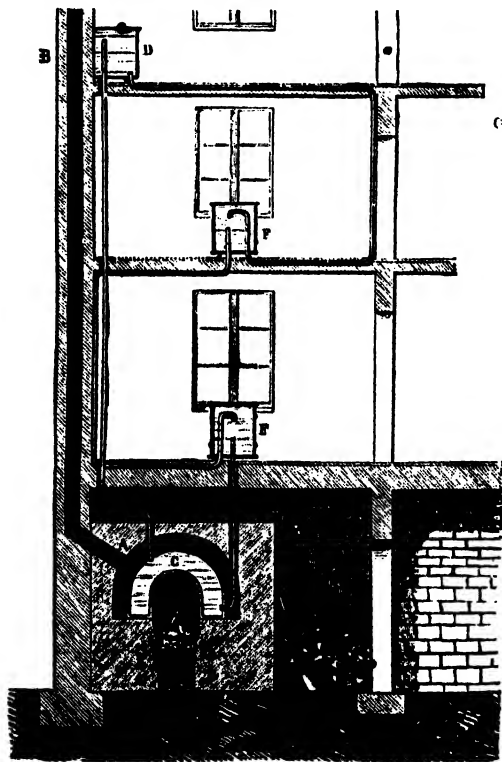


Fig. 44.—Heating a building with hot water: A, B, chimney ; C, boiler ; D, F, F, cisterns. The hot water rises from the boiler (C) through the straight pipe to the cistern (D), from which it returns by bent pipes passing through F, F, until it again enters the boiler (C).

that the housewife should know how the water, which, of course, is heated by the kitchen fire, finds its way to

the bath-room. In addition, churches and schools, and sometimes private houses, are now warmed by the agency of hot water, and not by fires. The same facts that explain the flow of water, upon heating, from the kitchen boiler to the bath-room will also account for the circulation of hot water through a system of pipes, so that the two things may be studied together.

### PRACTICAL WORK.

1. Put a piece of ice at the bottom of a long test-tube, and weight it down with a piece of lead. Add water. Hold the tube obliquely and boil the water over a flame, and observe that the ice does not melt.

From this experiment we learn that water (and the same applies to most liquids, except mercury) is a bad conductor of heat.

2. Place a little lime on the palm of the hand, and bring a red-hot poker to touch the lime. The hand is not burned, because the *air in the lime* does not conduct the heat. If a piece of solid lime, out of which the air has been driven, is held in the hand, in contact with the red-hot poker as before, the hand

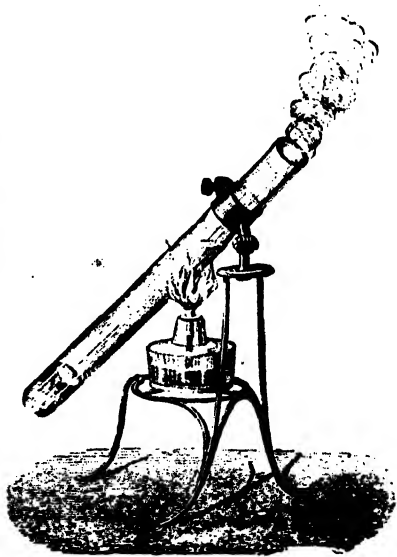


Fig. 45.—Water boiling over ice.

will now be burned, because lime in itself is a fairly good conductor of heat. Air—and other gases resemble it in this particular—is thus a very bad conductor of heat.

These two experiments have shown us that both liquids and gases are bad conductors of heat; yet heat is certainly transmitted through liquids and through solids, for the heat from the kitchen fire gets up into the bath-room and round the heated building, and the heat from the fire warms every part of the room. How this is brought about will appear from the following experiments:—

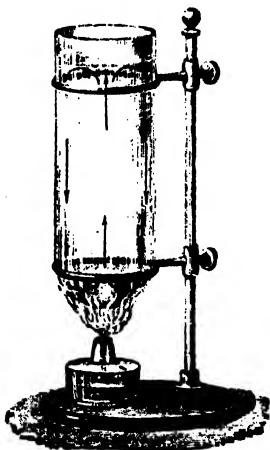


Fig. 46.—Currents produced in water by heat.

3. Heat over a small flame a flask full of water, into which put some solid colouring matter, such as cochineal, indigo, etc., or place within the water some sawdust. When the water becomes heated, currents will be seen by means

of the colouring matter; and it is evident that currents are set up in liquids by heat, whereby heat is carried throughout the substance of the liquid, although the liquid itself is a non-conductor of heat. The next experiment shows this to be true of gases also.

4. Fasten the end of a candle to a plate and light the candle. Surround the candle with water. Place over the candle a lamp-glass. The candle is slowly

extinguished, because the oxygen of the air inside the glass is soon used up, and no fresh air can get in.

Repeat the experiment, first introducing down the lamp-glass a diaphragm of cardboard, as in the diagram. The candle now continues to burn; for the effect of the diaphragm is to allow the used-up air to pass out by one side and for the pure air to enter by the other. A smoking match held in the chimney will demonstrate that there is an upward current in one half and a downward current in the other half of the divided chimney. This shows that the action of heat upon gases is to produce currents in them, by which, as in the case of water, the heat from the source is transmitted through the substance of the gas. In this way winds and the currents referred to under ventilation are produced, and in this way also the heated water is carried completely round a circuit of pipes, as in the heating of a building or in the transference of hot water from the kitchen boiler to the bath-room and back to the boiler again.

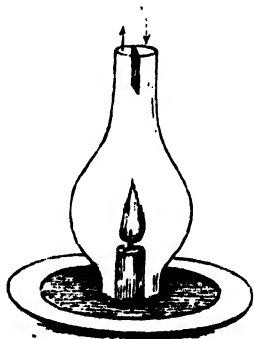


Fig. 47.—Currents produced in air by heat.

**Conduction, Convection, and Radiation.**—We may now proceed to enumerate the three ways in which heat may be transferred to places remote from the source of heat. Place a poker in the fire: the end away from the fire soon becomes visibly warmer than it was before the poker was placed in the fire. Next place a piece of wood resembling the poker in size

and form in the fire, and note that it burns, but no increase in the temperature of the end remote from the fire is observed. Hence heat travels easily through iron, but not through wood. Not only are the two ends of the poker made hot, but the whole of it is heated, for the heat travels through the poker by the successive change of all the parts of it from cold to heat.

This mode of heat transmission is known as **conduction**. Iron and all metals are good conductors of heat.

Through gases and liquids heat does not pass in this way. The particles near the source, being heated, expand, and so become lighter bulk for bulk than the other particles. They rise, and transmit their heat to the other particles by contact with them. To prove this, take a flask in which is placed some water and some solid colouring matter such as cochineal. Heat the water, and note the currents that are set up, starting from the water near the source of heat. This mode of transmission, in consequence of the actual movement of the heated particles, is known as **convection**, or carrying, and it is the way in which heat is transmitted through liquids and gases. Winds, ocean currents, and the heating of buildings by means of hot water are, as above noted, familiar instances of the working of this principle.

There is yet another mode of transmission, known as **radiation**, which is not so easy to explain to young students. By radiation, heat from a source, as from a fire or from the sun, travels in straight lines without any heating of the medium through which the heat passes.

When we stand in front of a fire, we are warmed

by radiant heat, since we feel the warmth, however cold the air may be.

On the top of a high mountain the sun's rays may feel uncomfortably hot, whilst the air itself is strikingly cold. This is because the sun's heat is radiant heat, and does not warm the air, except when the air is moist, which is not the case with mountain air.

In like manner, a light shining through the chink of a shutter is only seen when the eye is in a straight line with the light, since light, like heat, radiates out in straight lines from the source.

It will now be understood why we say that the yard is the distance between the two scratches on the golden plugs only when the temperature of the bar is  $62^{\circ}$  F., and why also we define the gram as the weight of a cubic centimetre of water only when the temperature of the water is  $4^{\circ}$  C. At temperatures above  $62^{\circ}$  F. the distance between the scratches would be more than a yard, and at temperatures above  $4^{\circ}$  C. the cubic centimetre of water would weigh less than a gram, on account of the expansion inseparable from heating.

## XI. CHANGE OF STATE AND LATENT HEAT.

**Cohesion.**—Matter exists in three states—namely, the solid, the liquid, and the gaseous. The reason given for this is as follows: there is a mutual attractive force, acting between particles of matter of the same kind, when such particles are in close contact. To its action are due the solidity and liquidity of solids and liquids respectively, the attractive force being greatest in solids, less in liquids, and non-

existent in gases. If blacklead is powdered, and the powder is subjected to great pressure, the particles cohere to form a solid; indeed, this is exactly what takes place in the preparation of blacklead for pencils. And so, generally, when the particles of a powder are brought into very close contact, they cohere, because the attractive force of cohesion is then able to assert itself. That water and other liquids form drops is an effect of cohesion; and the larger the drop the greater the cohesive force: thus mercury drops are larger than those of water, because the force of cohesion is stronger in the case of mercury than it is in that of water.

**Change of State.**—The conversion of solid ice into liquid water, and of this into gaseous steam, has been already referred to. Now there is very good reason for believing that this change of state from solid to liquid, and from liquid to gas, simply consists in overcoming the force of cohesion, and in driving the individual molecules composing the body further and further asunder, thus increasing the inter-molecular spaces, to which is due the porosity of the body. Each body has its own particular temperature for fusing or melting, and for vaporizing or gasifying. Thus, at the ordinary pressure of the atmosphere, the melting point of ice is  $0^{\circ}$  C. ( $32^{\circ}$  F.), and the boiling point of water is  $100^{\circ}$  C. ( $212^{\circ}$  F.). This we saw when graduating the thermometer. To determine the melting points and boiling points of substances is a very important operation. Expansion on heating and change of state are thus seen to be closely related; but the expansion in passing from one state to another is usually very great, and at times tremendously so. Thus steam occupies 1,700 times as much space as the water from

which it is formed; hence the explosive power of steam.

Water is an exception to this rule of expansion on heating, since ice upon melting contracts, and the water resulting from the fusion also continues to contract until a temperature of  $4^{\circ}\text{C}$ . is arrived at, which is the point of greatest density of water. Conversely, water expands on freezing, which explains why it is that ice floats on water.

### PRACTICAL WORK.

1. Insert a thermometer into masses of melted butter, lard, bees'-wax, sugar, etc., and find the point of fusion or melting point.

2. Fill a beaker with pieces of ice, and bury a thermometer bulb and as much of the stem as possible in the ice. Now apply heat, and notice that, no matter how much heat has been applied, the thermometer registers no increase in temperature until all the ice at  $0^{\circ}\text{C}$ . has been converted into water at  $0^{\circ}\text{C}$ .

3. Mix fifty grains of water at  $60^{\circ}\text{C}$ . with the same weight of ice at  $0^{\circ}\text{C}$ . Observe that only part of the ice melts, and that the temperature of the mixture is now  $0^{\circ}\text{C}$ .

4. Make two cold saturated solutions of Glauber's salt (sodium sulphate) and "hypo" (sodium hyposulphite). Insert a thermometer in each mass, and note that upon the solidification of the solutions heat is evolved, as is



Fig. 48.  
Latent heat.



proved by increase in the height of the mercury in the thermometers.

5. Take fifty grams of water at the ordinary temperature of the room, and pass steam into it from boiling water until a thermometer placed in the water registers  $100^{\circ}$  C. Weigh the water again, and notice how many grams of steam have been required. Thence calculate through how many degrees one gram of steam would raise one gram of water.

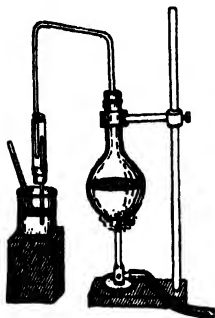


Fig. 49.  
Latent heat of steam.

It remains to explain these various facts. It will be noticed that there appears to be a loss of heat in passing from solid to

liquid, since, in the case of the melting of ice, the thermometer registered no increase in temperature, although there was a steady application of heat. Heat, as before stated, is a form of power or energy. When, therefore, heat is applied to a body in the process of liquefaction or of evaporation, we are adding power or energy to that body, which makes itself evident, not as heat, but as work, the particular work being the changing of the state of the body. This work is twofold in its nature. (1.) The attractive power of cohesion is either partially overcome, as in turning a solid into a liquid, or entirely, as when a gas is the result of the application of this power or energy. (2.) The pressure of the atmosphere is either partially or entirely overcome. It is again to be observed that this heat, which is employed in

• overcoming cohesion and in increasing molecular vibration, ceases to make itself sensible as heat. In consequence of this, the heat of liquefaction and vaporization is called **Latent** or hidden **Heat**. As seen from experiment, this latent heat again becomes sensible when a gas assumes the liquid condition, and when a liquid is reconverted into a solid. This explains why, when water-gas is changed into rain, heat is produced—that is, that rainy weather is often muggy weather. The latent heat of water is the highest: this is evident from the fact that it takes more heat to convert a given weight of ice into water, or water into steam, than in the case of any other body.

## **XII. PURIFICATION AND SEPARATION OF LIQUIDS.**

• We have just observed that to each substance corresponds a particular temperature for melting or fusing, and for vaporizing or gasifying. Sometimes the boiling points of liquids are very far apart, so that even a mixture of liquids may be separated by boiling, for the liquid with the low melting point will begin to gasify long before ebullition begins in the liquid with the higher boiling point. If this is true with respect to a liquid dissolved in or mixed with another liquid, it is much more true in the case of solids dissolved in liquids, for the solid will naturally have a higher boiling point than the liquid in which it is dissolved.

In the case of the separation of a more volatile from a less volatile liquid it is obvious that two different

operations will have to be performed—(1) vaporization, or conversion of the liquid into a gas; (2) condensation, or reversion of the newly-formed gas into a liquid. Thus a *boiler* and a *cooler* will be required; and the special name *still* has been given to this combination of a boiler and a cooler. The double process of vaporization followed by condensation is called *distillation*; and a most important process it is. The apparatus generally used in a laboratory for the boiler is a glass retort; and as the downward conversion from gas to

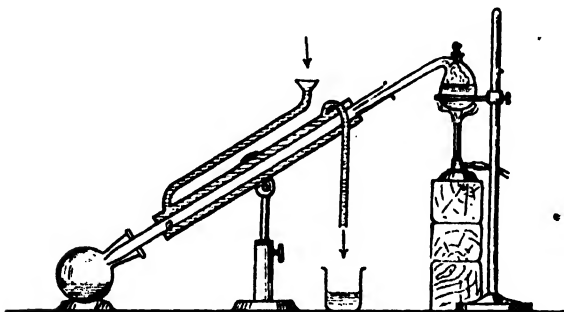


Fig. 50.—Apparatus used in distillation.

liquid is due to cooling, the cooler or receiver is kept at as low a temperature as possible. This is done by either running the tube connecting the boiler and the receiver through a tube which is arranged so that a current of cold water continually flows through it; or, more simply, a piece of moistened blotting-paper is placed on the receiver, by dripping cold water upon which the temperature of the cooler is kept down low enough for the incoming vapour or gas to liquefy. (Fig. 50.)

### PRACTICAL WORK.

1. Take solutions, containing ink, sugar, salt, milk, and separately distil them, whereby the solids may be separated from the liquid which comes over and condenses in the cooled receiver.

2. With the same apparatus, distil a solution of methylated spirit in water. Keep the gas or spirit-flame low down, so as to produce only a moderate amount of heat; then take the density of the sweet-smelling liquid that collects in the receiver. It is a mixture of alcohol and water; for some of the water is carried over bodily by the alcohol, and many distillations would have to be performed before even a fairly pure specimen of spirit could be obtained.

### XIII. MOISTURE IN THE ATMOSPHERE.

The evaporation of which we spoke in our last lesson does not take place only upon boiling a liquid. It is always in operation wherever the free surface of a liquid is exposed; indeed, it is not limited to liquids, but takes place in solids, as in the case of snow, which, in times of continued frost, will diminish in amount, and even disappear solely, in consequence of evaporation.

But evaporation of water when exposed to the atmosphere is the special case that we have to study, and we can best start from washing-day. How is it that the clothes dry when hanging out in the air? Simply because the moisture that is in them slowly evaporates. When does drying proceed slowly? On dull, damp, foggy days, and for this reason: the

amount of the evaporation depends upon the condition of the air with respect to moisture. Only a fixed quantity of moisture can be retained by the air; and when this amount has been taken up, evaporation, if not altogether stopped, is so slow as to be imperceptible. The drier the air, the more rapid is evaporation; and, again, the warmer the air, the more rapid the evaporation, for warm air can retain more moisture than colder air. Dew, cloud, mist, fog, rain, snow, and hail are all due to the deposition of moisture consequent upon the cooling

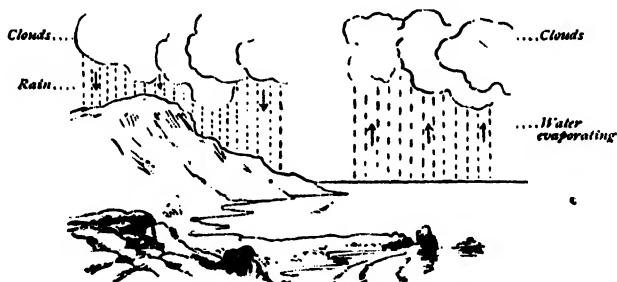


Fig. 51.—Evaporation.

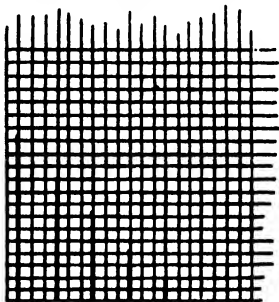
down of the air. As moisture in the gaseous condition is lighter than air, a mixture of air and moisture is lighter than dry air. This explains the fact that when the air is heavily charged with moisture the column of mercury in the barometer, which is supported by the air, becomes shorter—that is, the barometer falls. Hence a falling barometer indicates rain, simply because it tells of a lighter air—lighter on account of the excess of moisture contained in it. This moisture, upon cooling, from any cause whatever, will most likely descend as rain.

## PRACTICAL WORK.

1. Get a shallow vessel of very thin glass, so that the vessel may be as light as possible. A clock-glass used by chemical students will do very nicely. Place in the clock-glass a layer of water, and expose the vessel to the atmosphere. Weigh each day, and record the loss due to evaporation in graphic form (Fig. 52).

2. Weigh, day by day, a bag of seaweed or a roll of flannel, and record increases or decreases in weight, at the same time noting carefully what sort of weather is connected with the changes in weight.

The changes will, of course, depend upon the amount of moisture present in the atmosphere. Thus a great increase in weight will mean a very moist condition of the atmosphere, and *vice versa*. It is for this reason that a piece of seaweed is frequently to be seen hanging up in the halls of seaside houses. It acts as a rough kind of barometer.



Each mm. = 1 gm.

Fig. 52.

#### XIV. MOISTURE IN THE ATMOSPHERE.

(Continued.)

In order to determine more accurately the condition of the atmosphere with respect to the amount of moisture it contains, certain instruments, called hy-

grometers have been constructed. The simplest of all, and the one most commonly used, consists of two ordinary thermometers, on which account the instrument is described as the *wet and dry bulb thermometer*. From the diagram which is given below, it will be seen that the two thermometers are fixed side by side on a stand, and that one of them has its bulb covered with muslin, which is further connected

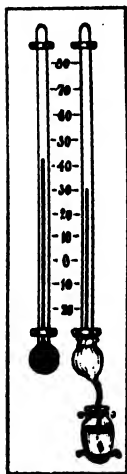


Fig. 53.  
Wet and dry bulb  
thermometer.

with a vessel containing water by means of a piece of wick such as is used in a spirit-lamp. Its action may be explained thus: the muslin is first of all saturated with moisture, and a further supply of moisture is carried to the muslin by the wick from the vessel, just as the petroleum is carried from the lamp reservoir to the burner by the wick—by capillary attraction, that is, by being sucked up the little passages that are in the porous wick.

Evaporation from this muslin will take place more or less rapidly, according as the air is dry or the reverse. But evaporation is change of state from liquid to gas, and this change requires power or energy in the form of heat to accomplish it. The required heat is taken from the muslin and from the bulb which it surrounds, and so the cooling of the covered bulb goes on so long as the air near the bulb can take up more moisture. The amount of the cooling can easily be seen from a glance at the wet bulb thermometer; and further, when the mercury in this no longer descends, but is

stationary, we know that the air in the vicinity of the wet bulb has taken up as much moisture as it can—that is, it is saturated.

Hence we may form an idea of the hygrometric condition of the air by means of this simple instrument; for when there is a considerable lowering of the temperature of the wet-bulb thermometer as compared with the temperature of the dry-bulb thermometer—that is, when the difference between the two thermometers is great—the air is very dry, and *vice versa*. Special tables have been constructed which enable us to determine the dew point, or to what temperature the air must be lowered before dew will be deposited.

#### PRACTICAL WORK.

1. Place a tumbler containing cold water in a warm room, and observe the deposition of dew on the cold sides.

2. Make a wet and dry bulb thermometer by taking the bulb-tubes from two ordinary thermometers and fixing them up as in Fig. 53. Take the readings of the two thermometers from day to day.

3. Take the thermometer and barometer readings daily, and make a weather chart. Compare this with that given in a daily paper. Make a curve of pressures and of temperatures, as done in the newspaper that you consult.

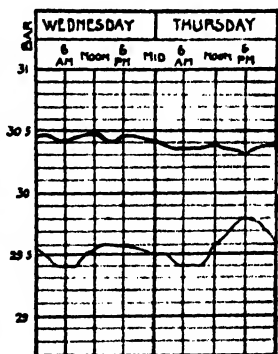


Fig. 54.—Weather chart.



**XV. SOLUTION, SOLUBILITY, MELTING.**

If you fill a cup completely full with tea, and then carefully place in the tea some moist sugar, bit by bit the sugar will disappear, yet the tea will not run over. How is this to be explained? If the cup was quite full to begin with, why was it not more than full when the sugar was added? The answer given is that the molecules of the water must have been some distance apart, and that the sugar has been deposited between the various molecules. Thus liquids as well as solids and gases are all porous, and solution may be explained as the process whereby the pores of a liquid receive the particles of a dissolved body. Water has been called the *universal solvent*; and although this is not strictly true in its literal sense—as there are many substances, such as pebbles, sand (silica), and heavy spar (barium sulphate), upon which water has no solvent action—still, as pointing out the property that water possesses of dissolving much of many substances, and little of many more, the term is a very useful one. Indeed, so great and so varied is the solvent power of water, that it is questionable whether absolutely pure water has ever been looked upon.

Salt, methylated spirit, and ammonia, are examples respectively of a solid, a liquid, and a gas soluble in water.

**Action of Heat.**—The action of heat upon solution is very interesting. If powdered nitre, common salt, Epsom salts, and potassium chlorate are severally dissolved in cold water in small instalments until an added portion is no longer dissolved, it will be seen

that the solvent power of water varies greatly for different soluble substances. When no more can be dissolved, the cold solution is said to be **saturated**. If heat is now applied, it will at once be evident that the **solvent power of hot water for solids is greater than that of cold water**; for further additions of the powdered substances to the solutions are readily and completely dissolved, and considerable amounts may be added before a **hot saturated solution** is obtained.

As regards liquids, the solvent power of water is much the same as with solids, as it increases with additional heat; but with gases the conditions are exactly opposite. Thus, if heat is applied to the solution of ammonia or spirit of salt (hydrochloric acid) in water, gaseous ammonia and hydrochloric acid are evolved, thus proving that **the solvent power of liquids for gases decreases with an increase of temperature**.

**Pure and Hard Waters.**—The difficulty of obtaining pure water has already been referred to. **Rain-water is the nearest approach to pure water to be found in nature**. Yet this, in falling through the atmosphere, dissolves carbon dioxide and other soluble matters in very slight degree. Spring-water, as everybody knows, is charged with various substances that have been taken up by the water in its passage through the earth, **which dissolved substances are the cause of the hardness of water**.

One of the commonest substances taken up in this way by spring-water all the world over is **limestone**; and it is very interesting to know **how spring-water dissolves limestone rocks**, for pure water is incapable of dissolving this common rock substance. Water, however, charged with carbon dioxide very freely dissolves

limestone. When, therefore, in its passage through the atmosphere, or through underground channels, water takes up carbon dioxide, it is increasing its solvent power for limestone, of which large quantities are thus held in solution by spring-water. When, either by heating or by decrease of pressure, the carbon dioxide is released, the limestone is deposited, as in the case of the fur of the kettle, the scale of the boiler, and the stalactites and stalagmites of fairy grottos.

By **boiling**, or by adding lime or soap, such water is softened. The boiling drives off the carbon dioxide, whilst the lime combines with this substance, the result being in each case that the water, no longer containing carbon dioxide, is incapable of holding the limestone in solution. Soap forms with the limestone dissolved in water an insoluble substance that floats on the water, as may be seen when we wash ourselves, and a soluble body that tends rather to soften than to harden the water. Bodies that are insoluble in water may readily dissolve in other substances, as will be evident from what follows.

**Melting** and **dissolving** are two very different things, the former being a simple liquefaction by heating, the latter a mixing of a solid (liquid or gas) with a liquid (solid or gas), whereby liquefaction results. Solution acts without heat, whereas fusion requires the application of heat, the solid being re-formed on cooling.

### PRACTICAL WORK.

1. Fill a wine-glass quite full with water, and carefully drop in some finely-powdered salt or sugar. Observe that the water does not overflow. •
2. Place a measured volume of some ordinary tap-

water in a watch or clock glass, and carefully evaporate to dryness. Weigh the residue. \*

3. Test the solubility in water of as many substances as possible—for example, salt, sugar, nitre, washing-soda, chalk, hearthstone, marble, etc. Arrange these bodies in the order of solubility, distinguishing clearly between those that are soluble and insoluble respectively in water.

4. Test the solubility of the latter of these two classes of bodies—such as iron, hearthstone, chalk, and marble—in vinegar and in spirit of salt (hydrochloric acid).

5. To water add powdered salt, sugar, soda, nitre, borax, alum, lime, and chalk, until no more can be taken up by the water. Such a solution is called a **saturated solution**. Now measure out ten cubic centimetres of each solution on to a counterpoised clock-glass, and after evaporating to dryness ascertain the weight of each solid contained in ten cubic centimetres of each solution. Record these results graphically on squared paper.

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## **XVI. TRANSMISSION OF HEAT: CLOTHING.**

Our previous work has shown us that heat is transmitted—(1) by conduction through the substance of conductors, of which the metals are the best; (2) by convection, which is the way in which heat travels through liquids and gases; and (3) by radiation.

Our experiment with lime proved air to be a bad conductor of heat, and the boiling of water without melting the ice at the bottom of the tube illustrated the same fact with respect to water. These experi-

ments should be repeated until it is well understood that liquids, as a rule, are worse conductors of heat than solids, and that gases are the worst conductors of all ; for water and air may in this respect be taken as typical of liquids and gases generally.

If the hand is placed on a metallic surface, the sensation of cold is experienced, for the reason that these substances conduct heat readily away from the hand ; and it is this loss of heat that is termed cold. For the same reason, substances warm to the touch, such as flannel, wool, feathers, and furs, are bad conductors of heat, on account of which they are used for clothing and for bedding. This explains why the ice merchant wraps his ice in blankets, which keep the heat away from it by virtue of their low power of conductivity for heat.

#### PRACTICAL WORK.

1. Repeat experiment in proof of the low conducting power of air.

2. Repeat experiment in proof of the low conducting power of water.

3. Repeat experiment in order to again show convection currents in boiling water, etc.

4. Repeat experiment to show that convection currents are set up by heat. Observe the chimney of a fire grate, and by means of the smoke from a piece of smouldering brown paper, show the existence of the same convection currents in the neighbourhood of the flame of a candle, lamp, or gas-burner.

5. Stand with your back to a fire and take the temperature of the air with a thermometer. Now let the fire play upon the thermometer, which, be it

observed, is now no nearer than at first. Note the increase in temperature, which is due to heat that has been transmitted by radiation.

6. Examine various clothing materials, and compare the warmth of the feeling when touching them, remembering that those feeling warmest are the worst conductors. Arrange them in a list according to their conducting power as thus ascertained.

7. Write an account of how the body is kept warm in the winter and cool in the summer by non-conducting clothing.



## Part II.

### I. COOKING.

**Heat and Food.**—Just as the steam-engine continually requires fuel if it is to continue working, so also the body must be supplied with food if it is to maintain the performance of its functions. Without food the body cannot grow; without food there can be no replacement of the parts destroyed by the wear and tear of life; and without food there can be no force to keep the body going. We must eat to live.

**Digestion.**—Eating, if it is to be of any use, must be something more than the mere taking of food material into the body. The food must be digested—

that is, it must be broken up and dissolved in such a way as to render it capable of easy passage into the blood stream of which it is to form a part. To bring about this food breaking up,



Fig. 1.—The digestive organs.



or digestion, is the special work of **mouth, stomach, and intestines**; and, under the most favourable circumstances, this work of digestion makes great demands upon these organs and upon our bodily force. Hence it behoves us to strive to make digestion as easy as possible, by taking care to choose the **best forms of food**, and by doing as much as possible towards the **digestion** of these foods, **before they enter the body**, as may be. And we can do very much indeed in the way of preliminary or **outside digestion**, as will appear from what follows.

**Digestion begins in the Kitchen.**—That this is true is apparent from the fact that it is much more difficult to digest a piece of meat when raw than when cooked. And so of foods other than meat. Cooking, then, and especially **good cooking**, is really the **first stage in digestion** or blood-making; and if this be true, we must place the good cook amongst the number of our **best friends**.

**The Cook's Good Fairy.**—By a good cook a good fire is a thing greatly to be desired, for heat, under some shape or form, is the good fairy of cookery. It is heat that makes food **softer by breaking up fibres and cells**. It is heat that **kills harmful parasites** that haunt both vegetable and animal food. It is heat that **prevents food from putrefying**, and that makes it **agreeable to the taste, pleasant to the smell, and comely to the sight**, whereby the jaded worker or the appetiteless invalid may be incited to take that food, which they need all the more because they crave for it the less. With much greater force, then, can it be said of fire than of wine, that it "**giveth a cheerful countenance, and maketh glad the heart of man.**"

**Both Radiant Heat and Heat from Contact are used.**—

An examination of the several ways in which heat is applied to food, in order to cook it, will show that such heat is obtained both by radiation from heated bodies and by contact with them. In **roasting, baking, and grilling**, the food does not come into contact with the fire, but is cooked by the heat that radiates therefrom, whereas in **boiling, frying, and steaming**, the necessary heat is got by contact with the heated water, oil, or steam. How heat may be made most effective in cooking will appear from the experiments that follow. Here it need only be said that there is much for the housewife to learn if she would be the mistress of fire in respect to cooking, for here, as elsewhere, fire is a good servant but a bad master.

**Different Kinds of Food.**—Before proceeding with our experiments, it will be convenient to say something about the various kinds of food, the action of heat upon which we are about to study.

**1. Mineral and Organic.**—The simplest way of dividing foods is that of considering them as consisting of things that have and of things that have not had life. Common salt and water will belong to the former class; beef and potatoes to the latter class, which may be further divided into **animal and vegetable foods**.

**2. Heat-giving and Flesh-forming.**—This is a very old mode of division, and it is important because it clearly points out the chief uses of food—namely, (1) to build up the body; (2) to supply it with heat and force. Unfortunately, however, it is not possible to have a clear division on these lines, as the **flesh-forming or nitrogenous foods**, as they are called, such as

lean of meat, white of egg, gluten of flour, etc., that contain nitrogen, and usually build up the tissues of the body, do not always do this. On the other hand, fats, starch, and sugar, which contain no nitrogen, but usually are burnt up slowly in the body, and so produce heat and force, sometimes are stored up, and so build up tissue rather than supply heat and force from their combustion.

3. **The best Division.**—For this reason it is best to enumerate five divisions of food without particular reference to what they do, remembering, however, that all food is concerned either in building up tissue or in producing the heat and force necessary to life and movement of any kind.

#### FIVE KINDS OF FOODS.

Nitrogenous.	Fats.	Carbohydrates.	Salts.	Water.
Gelatine, from bones. Albumen, in white of egg. Myosin, in lean meat. Legumen, in peas and beans. Gluten, in bread. Casein, in cheese.	Olein, } Palmitin, } in Stearin, } fats Butyrin, } and } oils. } in } butter.	Glucose, in grape sugar. Sucrose, in cane sugar. Lactose, in milk. Starch, in flour, rice, potatoes, and vegetables generally.	Common salt, vegetable acids, etc.	1. That contained in all foods. 2. That taken as a beverage, either as water, or in the form of alcoholic drinks, mineral waters, etc.

#### PRACTICAL WORK.

1. Break an egg and pour the white of it into a test tube in which a thermometer is placed. Next insert the test tube into some water and slowly

heat it. Observe the white of the egg when the thermometer registers a temperature of  $135^{\circ}$  F.: it has begun to thicken or **coagulate**. Now increase the temperature to  $180^{\circ}$  F., and notice that the liquid in the test tube becomes jelly-like in consistency. Finally, boil the water, and the white of the egg completely coagulates and becomes white and opaque.

**Albumen.**—This substance, which upon the application of heat changes from a watery liquid to a tough, horny, opaque solid, is albumen, one of the most important of nitrogenous foods. Albumen consists of the elements carbon, oxygen, hydrogen, nitrogen, and sulphur. It is especially valuable as a flesh-forming food. In the liquid condition, as in an uncooked egg, it is easily soluble in water; but in the opaque, horny form, as seen in a hard-boiled egg, it does not dissolve in water at all.

Which of these three forms is best for digestion? The jelly-like form that obtains when the temperature is about  $160^{\circ}$  or  $180^{\circ}$  F., and this should be the condition of the albumen in a properly cooked egg.

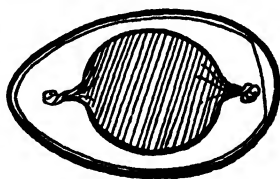


Fig. 2.—Section of egg.

**2. How to Cook an Egg.**—Boil some water in a saucepan, and then remove the saucepan from the fire. Place an egg in the water, and let it remain there for about five minutes, when, upon being cracked, the white will be found to be of the jelly-like consistency referred to above, in which condition it may be most easily digested. The cold egg reduces the temperature of the boiling water to about  $180^{\circ}$  F., at which tem-

perature the desired jelly-like condition of the albumen obtains.

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**Digestibility of an Egg.**—From observations made on a young Canadian, into whose stomach an opening had accidentally been made, it was found that whereas a soft-boiled egg was digested in three hours, a hard-boiled egg required three and a half hours for digestion.

**Application of these Experiments.**—The coagulation of albumen is of first-rate importance to the cook, since successful roasting, baking, grilling, and frying of meat depend upon its application. In cooking meat by any of these ways, it is essential that the nutritious meat juices should be kept within the joint. Now meat contains much albumen in the form of blood serum, and if the surface of the joint, immediately upon being put down for cooking, is raised to a temperature sufficiently high to coagulate this albumen—namely, to a temperature of from 160° to 180° F.—a coating of coagulated albumen is formed which effectually prevents the escape of the savoury juices, and so keeps the meat juicy and nutritious. Hence the fire, the oven, the water, or the fat in which, or by which meat is cooked, should be sufficiently hot to at once raise the outside of the meat to the temperature of coagulation. After this has been done, the gradual cooking of the meat may be proceeded with, there being now no risk of losing the juices, which are kept within the meat by the outward coating of coagulated albumen.

**Boiling for Soup.**—In soup-making our object is to extract the nutritious parts from the meat, for which reason we place the meat, cut up into pieces, into cold water, which is made to simmer, but is not allowed to

boil. There is thus no coagulation of the albumen, and therefore nothing to prevent the extraction of the meat-juices, on the presence of which the richness of the soup depends. It is thus to be observed that a knowledge of the action of heat upon albumen settles how we are to cook, both when we do, and also when we do not, wish to keep in the meat juices.

### PRACTICAL WORK.

3. Carefully weigh an egg, and then proceed to boil it. After carefully wiping off the water, re-weigh, and again boil until the egg is quite hard, when again it is to be weighed. Record your results. What may we infer from the various weighings?

4. **Action of Heat upon Fats.**—Cut up some mutton or beef suet, and place the pieces in boiling water contained in any convenient vessel. Observe the globules of fat that float upon the water. This shows that fat soon melts; a fact that is also shown when the heated suet is squeezed, by which means the membrane or skin is readily separated from the fat—now liquid—contained within it.

5. By melting some mutton fat, some beef fat, and some bacon fat in different test tubes or beakers, ascertain by the aid of a thermometer the melting points of the three varieties of fat, and record your results. Then let each of the three liquid fats solidify, and note the temperature of solidification, and again record your results. This affords a simple, if rough, test of the digestibility of various kinds of fat, for the fat that solidifies most easily, after cooking, is the least digestible. The order of digestibility should be (1) bacon, (2) beef, (3) mutton.

6. Take a piece of cheese, which consists largely of casein, a very valuable nitrogenous food-stuff, containing twice as much nitrogenous or flesh-forming material as meat. Toast the cheese, and observe how tenacious it has become. As food has to be broken up prior to digestion, it is easy to see how difficult of digestion toasted cheese must be. Indeed, cheese at the best is not easily digested, and should only be eaten in quantity by persons of robust digestion.

## II. ACTION OF HEAT ON VEGETABLE FOODS.

Before proceeding to experiment on vegetable foods, we must pause to inquire into the nature of starch, which is the most important constituent of all vegetable foods.



Fig. 3.—Starch granules.

**Starch.**—Potatoes, wheat, rice, sago, and arrowroot consist largely of a substance which, when examined under the microscope, exhibits the

peculiar appearance seen in the accompanying diagram. It is evidently granular; but the minute granules from different substances differ in appearance, so that by examining starch under the microscope it is possible to determine the body from which it has been obtained—that is, whether it is rice starch, potato starch, etc.

## PRACTICAL WORK.

1. **Preparation of Starch from Potatoes.**—Cut a potato in halves, and rub the two portions together,

letting a gentle stream of water drop upon the potato during the rubbing. Collect the water in a beaker, and observe that a deposit gathers at the bottom of the beaker. To show that this really is starch, take a solution of iodine in potassium iodide, and, having drained the water from the starch, proceed to treat it with a drop of the iodine solution. The fine indigo blue colour produced is a proof of the presence of starch, for iodine always stains starch an indigo blue.

**2. Action of Heat upon Starch.**—Make a thin cream with water and starch in an evaporating dish. Add twice the volume of boiling water, and mix well. Observe that the starch swells up and forms a paste, which will become quite liquid if sufficient water is present.

If this reaction is watched under the microscope, the granules of starch are observed to swell until the membrane of the granule bursts, when the contents of the little sacs escape, and are to a great extent dissolved in the water. As for the sacs themselves, they float about, and do not in any measure dissolve.

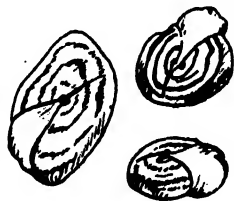


Fig. 4.—Bursting of wall of starch granules.

**Application of the Experiment.**—This experiment shows how necessary it is that all starchy foods should be boiled for a considerable time; for, before the starch can be available as food, the sac-wall must be burst, and if this has to be done in the stomach, so much more work is put upon that hard-worked organ. It is therefore wise and prudent that this part of starch digestion should be done in the kitchen.



3. **Further Changes in Starch.**—Into a large flask, capable of holding a litre, place 300 c.c. of starch solution prepared as above. Add with constant stirring 3 c.c. of dilute sulphuric acid. Heat to boiling, and maintain the temperature at boiling point for some time. With a glass tube withdraw some of the boiling contents of the tube at intervals, and test each separate portion of the liquid from time to time with iodine solution, until at length it no longer is coloured blue, but gives a reddish-brown colour when tested with the iodine solution. Now, as we have seen, starch always gives a blue coloration with iodine solution. The substance in the flask is **no longer starch**, for it does not do this. This new substance, that separates out as a flocculent precipitate when a portion of the boiled starch is poured into absolute alcohol—that is, alcohol free from water—is **dextrin**, or **British gum**, a specimen of which should be purchased and examined. It may be prepared in two other ways—namely, (1) by roasting the starch carefully, so as not to allow too high a temperature to result, for this would change the starch into **glucose**, as we shall see later, and (2) by the action of **diastase**, the substance produced in seeds when they germinate. British gum is used in calico printing, and also as a substitute for gum arabic, as on the backs of postage stamps.

4. **The Further Heating of Starch.**—Continue heating the solution in the flask, and constantly examine samples of it with iodine solution until no coloration whatever results. After about a quarter of an hour's boiling this stage will be arrived at. This further starch product is **glucose**, or **grape sugar**, a

substance of very great interest to us, because all starches must be converted into glucose before they can be taken into the blood.

5. Put a piece of ordinary bread into your mouth and masticate it for some time. Observe that, after a time, a sweet taste comes. This is due to the glucose that has been produced from starch by the **ptyalin** of the saliva, which, like diastase, has the power to change starch into the much-desired glucose.

### III. KITCHEN DIGESTION OF STARCHY FOODS.

**Bread Making.**—We have already spoken of the great service rendered to the digestive organs by heat in bursting the starch granules, and thus saving the stomach this labour. We can now see another way in which starch digestion goes on in the kitchen; for whenever starchy substances, such as flour, potatoes, rice, etc., are heated, some of the starch is changed into the sugary condition.

**Application to Bread.**—It will be easy to understand now why bread should be well baked, since the baking does the

two things above mentioned—that is, (1) it bursts the starch



Fig. 5.

granules, and exposes the starch to the action of the digestive juices; and (2) it changes some of the starch into glucose, which accounts for the sweetness of bread, as also for the indigestibility

of ill-baked, doughy bread. This, however, is only part of the work of the bread-maker, who must in addition make his bread **light and porous**. This is done (1) by **yeast**, (2) by **aeration**, (3) by **baking-powder**, concerning each of which a few facts may be experimentally learned.

### PRACTICAL WORK.

1. **The Action of Yeast.**—Fit up a flask, as in the annexed diagram, and place in the flask a warm solution

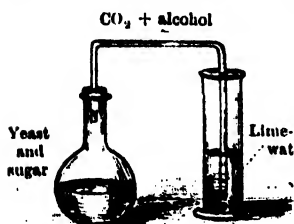


Fig. 6.—Carbon dioxide from yeast.

tion of grape sugar, in the proportion of one ounce of sugar to a gill of water, together with a piece of yeast as big as a marble. The temperature should be about 77° F. After a time

bubbles of gas collect in the tube, which gas, upon examination, is found to turn lime-water milky, and to answer all the tests for carbon dioxide. It is **carbon dioxide**. Furthermore, the solution in the flask smells strongly of alcohol, and it may easily be shown that it does really contain alcohol.

2. **Why Sugar and Yeast yield Carbon Dioxide and Alcohol.**—Yeast is really a very simple plant, which, when put into a sugary solution and made warm, begins at once to grow, and as a result of the growth of the yeast plant, carbon dioxide and alcohol are formed.

3. **To show that the Yeast Solution contains Alcohol.**—Pour some of the solution into the bulb of a retort, the beak of which should be led into a cooled flask, as

in the diagram. This combination of a boiler and a cooler is called a **still**. Apply heat to the bulb, and keep the receiver cool by placing blotting-paper upon it, cold water being from time to time poured upon the blotting-paper for the purpose of keeping it

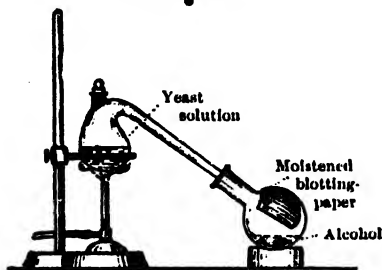


Fig. 7.—Alcohol from yeast solution.

cool. What takes place? The alcohol soon boils, and, passing over to the cooled flask, again condenses. If this process of **distillation** is repeated several times, whereby the flask contents be-

come more and more free of water, the characteristic smell of alcohol, or spirit, as it is commonly termed, will become quite strong. In addition, the liquid may be made to burn just as spirit does, and, in other ways, which at present you would be unable to understand, it may be shown that the liquid thus obtained by the action of yeast upon sugar is none other than alcohol.

**4. The Yeast Plant under the Microscope.**—Moisten a glass slip with some of the yeast solution, after the action has commenced. Place over this a cover glass, and proceed to examine the yeast solution under a microscope of moderate magnifying power. The little, oval, whitish bodies are the yeast plants, and the little lumps that appear on them, and after a time separate from them, are new plants, this being



Fig. 8.—Yeast plants.

the way in which these simple plants multiply. It is known as budding. When it is noticed how quickly these new cells come into existence, it cannot be longer matter for marvel that even a small quantity of leaven<sub>e</sub> or yeast should be able to leaven effectually the whole mass of the dough or sponge, however many times this latter may be greater in mass than the yeast itself.

5. **Fermentation and Ferments.**—Expose some milk to the atmosphere for a time, and note that it becomes sour. The reason for this is that a new body—namely, **lactic acid**—has been formed, and its formation is due to an exceedingly small organism present in the atmosphere which is able to feed upon, and grow in, milk, and, in growing, it changes the milk into lactic acid. Again, expose some beer, wine, or other alcoholic liquid to the atmosphere, and a new body—**vinegar or acetic acid**—results, which is also produced by the growth of a microscopic organism in the liquid.

Here, then, are three very common examples of chemical changes in organic liquids brought about by the presence of certain substances which do not themselves change. These chemical changes are what we mean by fermentation, and the substances by means of which the changes are induced are called **ferments**, of which there are two kinds—namely, **organic or living**, as in the three cases above cited, where the growth of the organism brings about the changes; and **inorganic, non-living, or soluble**, as in the case of the **ptyalin** of the saliva and the **pepsin** of the gastric juice, by the agency of which starch and cane sugar are changed

• into **grape sugar**, or **glucose**, in the mouth by **ptyalin**, and **proteid matters** are changed into **soluble peptones** in the stomach by **pepsin**.

It is to be noticed that we have used glucose in our fermentation experiment, and the question may fairly be asked why we did not use starch or cane sugar. The reply is that, before starch or cane sugar can be fermented, they must be **changed into glucose**, which preliminary change can, however, be brought about by the yeast cells. Still, as this preliminary change occupies time it is better to begin with the glucose as we have done; but the fact that starch and cane sugar, when left for a time exposed to the air, become changed in like manner as the glucose, is a sufficient proof, both of the ability of the yeast ferment to convert these substances into glucose, and then into carbon dioxide and alcohol, and also of the presence of yeast cells in the atmosphere.

• It is an interesting fact that when liquids liable to fermentation are shut off from the air by a plug of cotton wool they do not ferment, which is a proof that something contained in the air, and which cannot get through the cotton wool, is necessary to the fermentation. This something is yeast.

**Application to Bread.**—When flour is mixed with water, and yeast is added at the desired temperature—namely, from 70° to 100° F.—fermentation at once takes place, and carbon dioxide is evolved, which, in forcing its way from the interior of the sponge, as the dough is called, produces the required pores and ensures the desired lightness. At the same time, the alcohol, being very volatile, is driven off by the heat, so that the only result of the fermentation, as far as

the bread is concerned, is the porosity which, be it noted, would result from the escape of any gas from within the sponge, whether carbon dioxide or not. But as it happens that carbon dioxide has no effect, deleterious or otherwise, upon the dough, and as, moreover, it can be so readily made use of in more than one way, it is the gas always used to make the sponge porous.

**Another Way of using Carbon Dioxide in making Bread Porous.**—This gas is soluble in water, and, when under pressure, the gas dissolves in water in large quantities. When the pressure is relieved, and especially when heated, the gas readily escapes from the water. This is the principle underlying the production of the so-called **aerated bread**, in which the sponge is made, not with water pure and simple, but with water charged with carbon dioxide in the way just spoken of, so that, when the dough is heated, the gas escapes and pores result, just as in the case of bread mixed with yeast. One great recommendation in the case of bread made in this way is that the whole of the work is done by machinery, the bread never being touched by hands in the process of making.

**Baking-powder.**—Here, again, the result is the production of a porous, and therefore light, paste by the escape of a gas from within the substance of the dough during the baking, the gas again being carbon dioxide, but how it is produced we must spend a little time in explaining.

## PRACTICAL WORK.

**6. The Production of Carbon Dioxide.**—Fit up a flask with a thistle funnel and a delivery tube, or, if

available, use a two-necked Woulff's bottle, as figured in the accompanying diagram. Before corking, place within the flask some **chalk, whiting, hearthstone, limestone, marble, or washing-soda**, all of which substances agree in containing carbon dioxide in combination with an oxide of a metal, and are, on this account, called **carbonates**, from which the carbon dioxide can readily be set free either by **heat** or by **acids**. The delivery tube having been placed at the bottom of a dry gas jar, we proceed to pour down the thistle funnel some easily-procured acid, such as vinegar, spirit of salt, or hydrochloric acid, or vitriol, as sulphuric acid is called. The result is the evolution of carbon dioxide, which, being heavier than air, drops to the bottom of the gas jar and drives out the air step by step until at length the jar is full. To ascertain when this stage has been reached, a lighted taper is held at the mouth of the jar, and when it is extinguished the jar is full, for bodies will not burn in carbon dioxide as they do in air; hence the gas is called a **non-supporter of combustion**. Having filled a jar thus, place some lime-water in another jar, and proceed to pour the carbon dioxide into this lime-water, just as you would pour in water, then shake up. The lime-water at once indicates the presence of the carbon dioxide by becoming milky. This shows, at the same time, that the gas is heavier than air, otherwise how could it have been poured into the second jar? Pour

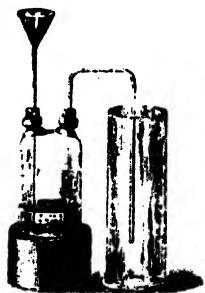


Fig. 9.—Preparing carbon dioxide.



another jar of the gas upon the flame of a candle. Why is the flame extinguished?

7. Put some solution of litmus in a jar, and let a current of carbon dioxide flow into it. Observe that the blue gradually changes to a purple tinge, which may be taken as a weak red. Hence, as acids turn red blue litmus, and other vegetable blues, such as that in pickled cabbage, we may learn from this experiment that carbon dioxide, when passed into water, not only dissolves in it, but forms with it a weak acid—weak because a purple and not a red colour results.



Fig. 10.—Pouring carbon dioxide from jar.

8. Above some lime-water contained in a gas jar burn a candle, as shown in the accompanying sketch.

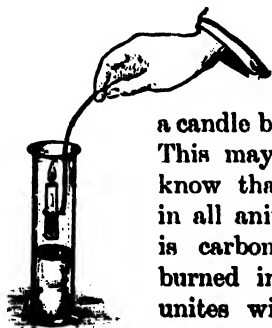


Fig. 11.—Candle burning over lime-water.

Shake up the lime-water, and observe the resultant milkiness, from which it may be inferred that, when

a candle burns, carbon dioxide is formed. This may easily be believed when we know that the chief element present in all animal and vegetable substances is carbon, and that when carbon is burned in air, the oxygen of the air unites with the carbon of the substance, the result of which combination is carbon dioxide. From this candle experiment, then, we infer that, when

any animal or vegetable substance is burned in air, carbon dioxide is produced; and it will be a useful

exercise for the student to prove this by burning as many such substances as possible, and by testing the products of the combustion with lime-water, whereby the truth of our statement may be established.

9. **How to make Baking-powder.**—We are now in a position to understand exactly what baking-powder is and what it does. Take of ground rice half a pound, of carbonate of soda one quarter of a pound, and of tartaric acid three ounces. Powder and dry the ingredients well before mixing them, and, after mixing, pass them through a colander or sieve. A serviceable baking-powder will now be obtained.

10. **Carbon dioxide from Baking-powder.**—Into a flask, fitted as before with leading-tube and thistle funnel, place some of the baking-powder as made in the last experiment. Pour some water down the thistle funnel, and gently heat the flask. What do you see? Test the gas that is coming off by passing the leading-tube into lime-water. What happens to the lime-water, and what does this show?

The essential parts of the mixture are the carbonate and the acid, for, as soon as these two substances are brought into close contact by being moistened, the acid decomposes the carbonate, just as in Experiment 6, p. 98, and as a result carbon dioxide is given off.

In MacDougall's baking-powder the two active ingredients are acid phosphate of lime and potassium carbonate. The principle involved is, however, exactly the same, there being present a carbonate and a solid acid, which, so long as they are kept dry, do not interact, but when water is added the acid acts upon and decomposes the carbonate, and so brings about the evolution of carbon dioxide.

**11. Composition of Flour.**—Make a little muslin bag and fill it with flour; then tie up the opening. Put the bag under a tap of running water, and whilst the water is playing on the bag squeeze the contents frequently, so that the starch may be quite carried away. When the bag is opened, a tenacious, sticky substance will be found within, which is not starch, since it does not turn blue when treated with iodine solution. If exposed to the moisture of air, the **gluten**, which is the name given to this sticky substance, will putrefy; but if kept perfectly dry, it remains unchanged as a glue-like substance.

**12. Gluten contains Nitrogen.**—Pound up together some lime and caustic soda. Put this mixture, together with some gluten, into a test tube, and heat over a Bunsen burner or spirit-lamp. Observe the strong smell of ammonia that comes off. This is due to the nitrogen and the hydrogen that the gluten contains; for these two elements combine to form ammonia when nitrogenous material, either vegetable or animal, decomposes, so that the production of ammonia, when a body is heated with soda lime, is the usual proof that such a body contains nitrogen. We may, therefore, class **gluten** amongst **nitrogenous** or **flesh-forming foods**, a fact that vouches for the justice of the claim of bread to be called the staff of life, since it contains **heat-giving starch** and **flesh-forming gluten**, the proportions of the two being starch 70 per cent., gluten 10 per cent.

**Action of Heat on Sugar.**—We have studied the action of heat upon starch, and have observed the bursting of the wall of the starch granule and the conversion of starch into British gum, or dextrin, and,

later on, into glucose, or grape sugar. The application of this to bread-making has led us to the study of the composition of flour. We now proceed to examine the effect of heat upon sugar, which, as a food-stuff, is no less important than it is pleasant.

13. Heat some white sugar in an iron spoon, or, better still, in a crucible. Find its melting point by inserting a thermometer in the melting mass, and record the result, which should be about  $160^{\circ}\text{C}$ . If now allowed to cool, we have the familiar **barley-sugar**, which differs from ordinary loaf-sugar in being yellow, glassy, and non-crystalline. If we continue the heating up to about  $200^{\circ}\text{C}$ ., we shall produce a dark-brown substance, the burnt-sugar of the kitchen, the proper name for which, however, is **caramel**; it is sugar deprived of some amount of water. Caramel is used for colouring gravy, and also for colouring wifes and spirits. Heated above  $200^{\circ}\text{C}$ ., the sugar is completely turned into a black mass of carbon.

14. That sugar contains carbon may be much more simply shown by taking a solution of sugar and adding thereto some vitriol or sulphuric acid, whereupon the liquid becomes black throughout in consequence of the production of a carbonized mass by the action of the sulphuric acid upon the sugar.

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#### IV. BURNING AND THE AIR.

##### PRACTICAL WORK.

1. Light a candle and place it on a piece of board floating on some water in a vessel. Place a glass bell jar, or a large glass jam jar, over the candle, so that

the mouth of the jar is under the water. The candle soon goes out, although there is still some air in the jar, else why does not the water rush up to fill the vacant space?

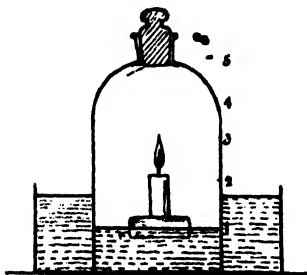


Fig. 12.—Candle burning under bell jar.

2. Using the same apparatus, but putting a piece of phosphorus floating on a cork in place of the candle and the piece of board, note how far the water goes up into the jar when the jar rests on the bottom of the vessel. Put a piece of

stamp-paper to mark the height of the water. Now take off the jar, and touch the phosphorus with a red-hot iron wire. The phosphorus is at once ignited, and burns brightly. Replace the jar, and press it down so that nothing may escape until the burning state has ceased. Let the apparatus remain as it is for about a quarter of an hour, when the white fumes formed by the burning phosphorus will have disappeared. Where have they gone? Evidently they must have dissolved in the water. Now mark with another piece of stamp-paper the height to which the water has risen inside the jar. If the experiment has been successful, this will be one-fifth of the jar, so that, in burning, the phosphorus has used up one-fifth of the air. It will be



Fig. 13.—Apparatus for obtaining nitrogen by burning phosphorus.

noted that there is still some phosphorus left, and the question now arises why the phosphorus and the candle did not go on burning. The only satisfactory answer to this question is, that there is something in air in which phosphorus, candle, etc., will burn, and something in which they will not burn; and that, whereas the former comprises one-fifth of the air, the latter comprises four-fifths. Taking this as the true explanation of the fact that the phosphorus and the candle only burn for a time and then go out, leaving four-fifths of the air unconsumed, it will be understood what is meant by the statement that **air consists of an active part which supports combustion, and an inactive portion in which things will not burn**, the inactive portion being four times the bulk or volume of the active part. The active portion of the air is called **oxygen**, whilst to the inactive part the name **nitrogen** has been given.

3. **When Bodies burn, new Bodies are formed.**—In our last experiment the white fumes that were formed, when the phosphorus burned, were neither phosphorus nor air, but a new body, named phosphorus oxide. This truth is rendered even more evident by the simple experiment outlined below. Take a piece of magnesium ribbon, and observe its colour, its lustre, and any other of its properties. Now insert the end of it in the Bunsen flame. It burns with a dazzling light, and, after the burning, there is left a white body, resembling neither magnesium nor oxygen. It is a new body, that has resulted from the combination of the magnesium with the oxygen. Whenever bodies unite together in the chemical way a new body results.

The student will perhaps ask what is meant by combination or joining together in the chemical way ; to which it must be replied that the study of chemistry is needed for a complete understanding of this.

Perhaps the following quotation from the author's "Section I. Physiography" will make matters a little clearer :—

"The gas nitrogen, which is the most plentiful constituent of the atmosphere, has no odour whatever, and is so sparingly soluble in water as to be practically insoluble therein.

"The same may be said of the gas hydrogen, one of the two elements found in water. If we merely mix these two gases together, we have a mixture which, like its component elements, is inodorous and insoluble.

"But if we cause electric sparks to pass through the mixture, we soon make acquaintance with the marvellous effects of a chemical union. The gas ammonia is produced, which is one of the most powerfully odorous of all bodies, and is so exceedingly soluble in water that one volume of water will dissolve more than a thousand volumes of ammonia. Whence comes the newly-got odour, and solubility ? Certainly not from the hydrogen, and, just as surely, not from the nitrogen. All we can say is that these properties, absent before and present after chemical union, have, in some way or other, been gained by the body in the act of combination.

"Again, hydrogen, carbon, and nitrogen are non-poisonous substances, and a mixture of them is equally innocuous. But let chemical attraction take place

between them, and a new body—hydrocyanic or prussic acid—is produced, and this is one of the most active and most deadly of all poisons. Such examples might be multiplied indefinitely; for it is the invariable rule that, **when chemical action takes place, a new body results**, differing in characteristics more or less widely from the original matter from which by chemical action it was formed. Bearing in mind these facts, it is no wonder that chemical combination has been spoken of as the nearest possible approach to a creative act. There is, of course, no creation of matter when combination takes place, but there is an abundant development of new and frequently startling properties, which result from the working of this mysterious chemical attraction.

“To the chemist the field of nature does not merely present an assemblage of animate and of inanimate bodies, of solids, liquids, and gases; it is his business to go deeper into the question, and to seek to learn **how many distinct kinds of simple substances there are** in existence, and what are the properties of each. By building up (**synthesis**) and by breaking down (**analysis**) he is bent on ascertaining what combinations are possible, and what is the nature of the new bodies resulting therefrom. In a word, **chemistry is an experimental science, which by varied analysis and synthesis seeks to discover the composition and properties of matter.**

“**Elements, Compounds, and Mixtures.**—When examining a body, the first question the chemist asks is, whether it is an element, a compound, or a mixture.

“**An Element is a body from which nothing has been got but itself.** Thus, we may do what we will to iron, sulphur, or oxygen, and, however much we may



seem to change them, from each we shall succeed in getting nothing but itself—from iron nothing but iron, whether solid, liquid, or gaseous; from sulphur, sulphur; and from oxygen, oxygen. Hence these three bodies are regarded as fundamentally simple and elementary. There are in nature about **seventy-five elements**, out of which, by combination, the whole universe is built up."

**A Compound is a body composed of two or more elements combined together chemically.** Thus, water and lime are compounds, the former of oxygen and hydrogen, the latter of calcium and oxygen. The white body produced by the burning of magnesium is also a compound, the elements in it being oxygen and magnesium; and the fact that the compound is formed so easily and with so much energy is a proof that the attraction between oxygen and magnesium is very great. It is very carefully to be noted that this attraction between the chemical elements that brings about combination varies very considerably, so that one element may attract and be attracted by the whole of the seventy-five elements, whereas another may have so little of attraction as to make it very difficult to bring it to combine with any element at all. Now, as the student will have guessed, this just expresses what was meant when we stated that air consisted of an active and an inactive portion, for oxygen is one of the most active of the chemical elements, whereas nitrogen is just the reverse; so that when bodies burn in air, or are acted upon by air, it is always safe to say that nitrogen has nothing to do with the matter, as it remains inactive and unchanged throughout the whole reaction.

**Burning is Chemical Combination.**—After what has been said and done, it is hardly necessary to say that, when bodies burn, chemical combination is taking place; and that, when the burning takes place in the atmosphere, the combination is between the oxygen of the air on the one hand and the burning body on the other, the nitrogen being purely passive. It is, however, necessary to say that there are **many degrees of burning or combination**. Sometimes, as in the case of phosphorus and magnesium, the burning is exceedingly energetic; sometimes, as in our own bodies, which are **always burning** so long as we are alive, the combustion is so slow and the combination so comparatively inert as to produce no light, although the heat produced by the combination is very plainly evident; sometimes, again, as in the case of the rusting of iron, the combination is so inert as to produce no evident heat. Still all these reactions result in the production of **new bodies**. Each is a case of chemical combination, whether we call it burning or combination or not; the difference is simply one of degree, and not of kind.

It will now be convenient for us to study the preparation and properties of oxygen, in order to realize from experience what is meant by combination and the formation of new bodies.

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## V. OXYGEN.

**Oxygen.**—This important constituent of the compound water and of the mixture air we now proceed to examine.

## PRACTICAL WORK.

1. Procure as much of the red powder—mercuric oxide, or red precipitate, as it is commonly called—as will cover a shilling, and heat it over a Bunsen burner in a hard glass test tube. Observe that the powder turns black on heating, but regains its colour upon cooling. This is a mere physical change, temporary, and productive of no new body. Continue the application of heat, and place a smouldering splinter of wood at the mouth of the tube. Soon the chip burns with great brilliance and bursts into flame. Evidently a new body has been produced, with properties different from those of mercuric oxide. This is oxygen. If now the sides of the tube be examined, a mirror will be seen to have formed upon it. This is mercury; so that, by means of heat, mercuric oxide has

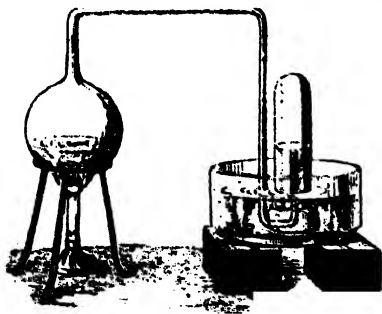


Fig. 14. — Priestley's experiment: oxygen obtained from air by the continued heating of mercury.

been broken up or decomposed into two new substances, oxygen and mercury, and all the signs of a chemical change are present.

**The Discoverer of Oxygen.** — This is an expensive and inconvenient means of preparing oxygen, but it is of exceeding interest, because it is the method by which Priestley, the discoverer of oxygen, prepared it, and demonstrated that it was a constituent of the atmos-

phere. He heated the metal mercury in air for three days in the apparatus shown in Fig. 14, and noticed that a red film formed upon the mercury. This was mercuric oxide, which had been obtained by the action of the air upon mercury. Subsequent heating at a higher temperature decomposed the body and set free oxygen, which for the first time was separated from air.

**2. Best Way to prepare Oxygen.**—Bend a piece of soft glass tubing to an obtuse angle at each end, and insert one end in a cork that tightly fits the mouth of a flask. Let the other end of the tube pass under a beehive shelf placed in a trough containing sufficient water to cover the shelf. Fix the flask on a sand bath resting on a retort stand. Take equal parts by weight of powdered chlorate of potash and manganese dioxide, with which the flask is to be about one-fourth filled. Over the hole of the beehive shelf place an inverted gas jar full of water. Apply heat to the flask by means of a Bunsen burner or a spirit-lamp placed under the sand bath. Soon bubbles of gas will collect in the gas jar, which, however, should not be placed on the shelf until after the escape of the first instalment from the delivery tube, as this will be air driven out by expansion and displacement.

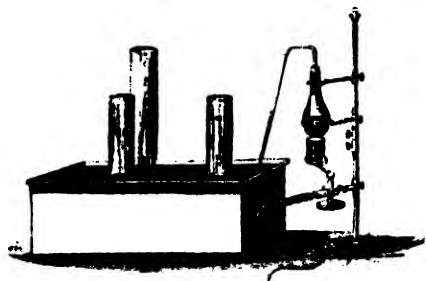


Fig. 15.—Apparatus for collecting oxygen over water.

Proceed now to collect eight jars of the gas, placing greased ground-glass plates over the top of the jars to prevent the gas from escaping.

**Jar 1.**—Place this again in water with its mouth downwards, and remove the glass plate. Shake the jar without allowing the mouth to come above the surface of the water, and observe that the level of the water does not show any perceptible increase in height. **Oxygen is almost insoluble in water.**



Fig. 16.—Smouldering chip in jar of oxygen.

**Jar 2.**—In this jar place a smouldering chip, as in the case of the gas that was obtained from mercuric oxide. Note that an exactly similar result is obtained, thus proving the identity of the gases obtained respectively from mercuric oxide and chlorate of potash.

**Jar 3.**—In a deflagrating spoon put a piece of sulphur about as big as a pea. Ignite the sulphur, and observe the pale blue flame with which it burns. Now place the burning sulphur in the jar of oxygen, and notice how much more brilliantly it burns. When the action is over, pour some water into the jar, and placing the hand over the mouth shake the contents briskly. Note that the hand is sucked down, thus proving that the sulphur oxide has been dissolved.



Fig. 17.—Phosphorus burning in oxygen.

**Jar 4.**—Treat this exactly the same as the preceding, substituting phosphorus for sulphur. Observe

that the phosphorus burns brilliantly in air, and with increased and dazzling brilliancy in oxygen. Note also that the oxide of phosphorus is even more soluble than the oxide of sulphur, as is proved by the hand being more forcibly sucked down. Keep the solutions thus obtained for further experiment later on.

**Jar 5.**—The deflagrating spoon is now to be filled with powdered charcoal, to ignite which will be more troublesome than in the two preceding cases. When the charcoal begins to burn, the deflagrating spoon is to be placed in the fifth jar of oxygen, as in the two previous experiments.

To each of the three solutions now proceed to add a solution of blue litmus, and note carefully that the sulphur and phosphorus oxide solutions at once turn the blue litmus a bright red, and that the oxide of carbon solution renders the blue purple.

If the two former solutions be tasted, it will be noticed that they are sour, like vinegar; moreover, they are like vinegar in turning vegetable blues red, as is shown by the action of vinegar on pickled cabbage. Such sour bodies that turn vegetable blues red are acids, whose properties we shall further consider later on. The point for the student to carefully note is, that an acid results when the oxides of sulphur and phosphorus are added to water, with which they therefore combine. With regard to the oxide of carbon, that also forms an acid when added to water; but the acid is a very weak one, as is suggested by the purpling instead of reddening of the blue litmus solution.

**Jar 6.**—Place in the deflagrating spoon a piece of sodium about as big as a pea. Heat it in the

Bunsen flame until it just begins to glow, then transfer the deflagrating spoon to the jar of oxygen. The sodium burns brightly. Add water to the white sodium oxide, and observe that it dissolves. Now take a portion of the liquid between the thumb and the forefinger, and notice the *soapy feel*. Pour into the solution some of the reddened litmus from one of the previous experiments, and observe that it is again turned blue; so that sodium oxide, when added to water, yields a soapy solution capable of turning red litmus blue, and of neutralizing an acid. Such a body is called an *alkali*.

**Jar 7.**—Put a coil of magnesium ribbon or wire in the deflagrating spoon, and hold the spoon in the Bunsen flame until the metal begins to burn with great brightness. Transfer the spoon to the jar of oxygen, whereupon the magnesium emits a blinding white light, and a white solid remains behind after the burning. Put some water in the jar, and observe that the oxide does not dissolve, as magnesium oxide is very sparingly soluble in water. Take now a strip of red litmus paper and moisten it. Place upon the moistened paper a portion of the white magnesium oxide, and note the blue stain. Magnesium oxide is a slightly alkaline body, like sodium oxide. Like sodium oxide, too, the oxide of magnesium is able to neutralize acids.

**Jar 8.**—In this jar we are about to burn iron. Take some fine iron wire and coil it up. Place it on the deflagrating spoon, and then sprinkle some sulphur on it. This is to start the action, as iron requires a somewhat high temperature before it will burn. After igniting the sulphur, immerse the deflagrating spoon

in the oxygen. At first only the sulphur burns; but by-and-by the iron begins to burn and to send out beautiful scintillations. If we look at the bottom of the jar, we shall see a deposit of a blackish substance that has burned its way into the glass. This is the oxide of iron. Upon the addition of water to it we find that there is no solution, as iron oxide is insoluble in water.



Fig. 18.—Iron burning in oxygen.

Our experiments have taught us that oxygen is a **very active body**, as it not only combines with metals and non-metals, but it does so with so much energy as to produce great heat and dazzling light. Instead of showing this with respect to six elements only, we might have multiplied our experiments very materially, as **oxygen combines with every known element except fluorine**, which is another way of expressing how extensively oxygen is attracted by and attracts other elements.

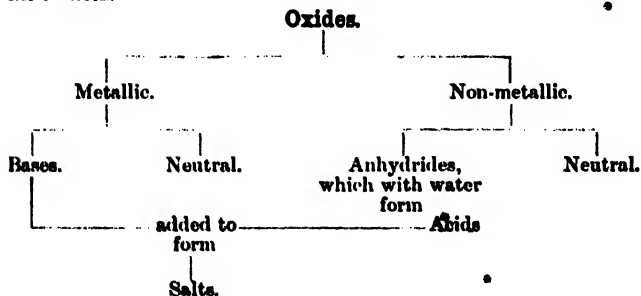
We have also shown that there are two classes of oxides which differ greatly in their properties, being, indeed, directly opposite in these respects; thus, sulphur, phosphorus, carbon, and other non-metallic elements, unite with oxygen to form acid-making oxides—that is, oxides which, when added to water, combine with it to form acids. These oxides are called **anhydrides**, a word signifying “without water”—that is, an oxide without the water needful to make it completely an acid.

Metals, on the other hand, form oxides mostly



insoluble in water—like oxide of iron—most of which, however, are capable of neutralizing acids. Such metallic oxides are called **bases**. Some bases—and they are the strongest, like oxide of sodium—are soluble in water, to which they give a soapy feel and an alkaline reaction—that is, ability to turn red litmus blue.

When an acid is neutralized by a base, a new body called a **salt** is formed; so that, from our study of oxygen, we have arrived at the meaning of three most important chemical terms—namely **acid**, **base**, and **salt**. It remains to be said that although the above statement is in a general sense true, there are some non-metallic oxides that are not anhydrides, and some metallic oxides that are not bases. The annexed table will make this clear, and will also indicate the intimate connection that exists between acids, bases, and salts.



**The Heating of Bodies.**—From what has preceded, it should now be thoroughly understood that to heat substances in air means to make the conditions favourable for the active part of air—namely, oxygen—to act upon these substances, and for these to

react upon oxygen, the end of this mutual action being the formation of one or more **new bodies** composed of oxygen and the substances in question.

3. In further proof of this, proceed to burn weighed portions of some or all of the following substances — fat and lean meat, bread, wood, paper — either in a weighed crucible or in a spiral of platinum wire. Weigh the incombustible ash that remains in each case, and write your weighings in your laboratory book. Where are the new bodies? They were gases, and have therefore vanished into air. Take a hard glass test tube, and place a piece of bread within it, and heat it. Observe the moisture that collects on the side of the tube. Evidently one of the new bodies is water. After having strongly heated the bread for some time, pour the gas produced into lime-water. Another of the gases is now seen to be carbon dioxide, and we learn that when such bodies as the above—that is, organic bodies—are burned in air, **water and carbon dioxide** are formed.

**Application to Ourselves.**—We have only to breathe on a cold surface, such as a pane of glass, to show that water is coming off continually from our bodies, even if we do not observe the water that comes off as perspiration; and again, we have but to breathe into lime-water to show that carbon dioxide is also constantly being produced within us. Something must, therefore, be taking place within us similar to the burning of meat, bread, etc. What is it? Of course it is our bodies which are slowly burning; and this accounts, not only for the water and carbon dioxide given off, but also for the heat which distinguishes a living from a dead animal.

4. **Heat without Air.**—Place a splinter of wood in a hard glass tube, and either plug up the mouth of the tube with cotton wool or hold the thumb loosely over it, by which means the air may be shut out. What do you observe? and what remains? You can see that when a body is heated out of air, the heat brings about the decomposition of the body, its volatile portions being driven off and its non-volatile portions alone remaining behind. In this case the fumes represent the volatile part, the charcoal the non-volatile part of the wood.

**Application to Coal.**—We shall see later on that exactly the same thing happens when coal-gas is manufactured from coal. Coal is placed in closed iron retorts and is strongly heated out of air, whereby it is decomposed, the volatile portion being impure coal-gas, the non-volatile constituent, corresponding to the charcoal of our last experiment, being coke, which remains behind in the retorts. Before being fit for use, the coal-gas has to be purified in various ways, of which we need speak no further.

Now let us return to our experiment. Take the charcoal from the tube and twist your platinum wire round it, so that you may be able to hold the charcoal in the flame. What happens now? What has become of the charcoal? What new body may have been formed?

**Application to Rotting or Decay.**—It may be shown that carbon dioxide and water are given off when organic matter, such as leaves, meat, etc., decays or rots in the air. If also the hand is plunged into a mass of decaying leaves, the mass is found to be warm. Does this not prove that decay is also slow

combustion or burning? Write down your reasons for this statement.

We may sum up the whole of what has been proved by saying that when animal or vegetable substances burn or decay in air, the chief part of what takes place is the action between oxygen, the active part of air, and the carbon and hydrogen, which are always present in such bodies, carbon dioxide and water being formed.

Do you understand now that life, death, and decay are mainly due to oxidation, and depend upon the active part of air? Think the matter over, and, if it is not clear, repeat the preceding experiments and read the preceding explanation of them.

## **VI. ACTION OF HEAT AND AIR ON MINERAL SUBSTANCES.**

We have already heated iron, sulphur, and phosphorus in oxygen, and we have found that iron combines or burns with difficulty in oxygen to form a black oxide; that sulphur readily combines with oxygen to form a suffocating gas that reminds us of burning matches. Why? Phosphorus unites with oxygen even more readily, burning as it does with dazzling brightness to form the very soluble phosphorous oxide.

It is now required to heat copper, iron, and sulphur in the air, in order to see whether the oxygen of the air will unite with them just as the pure oxygen did. Take, therefore, some copper turnings, and heat them in a test tube, or hold some copper wire in the Bunsen flame. It changes colour and becomes black, and the

student will readily guess that the black body produced is the oxide of copper. Now heat some iron wire, and observe that it becomes coated with a kind of scale, which is the oxide which has resulted from the burning. Sulphur is next to be heated in a test tube or crucible. Observe the pretty yellow liquid that is produced when the sulphur melts, and see how soon the suffocating odour and the blue flame become evident, thus plainly showing that oxygen and sulphur very readily combine together.

### PRACTICAL WORK.

1. **The Rusting of Iron.**—Fill a small flask or a test tube with water, and boil for a few minutes, by which means all the air dissolved in the water will be driven out. Now put a bright steel pen-nib in the water, and after having tightly corked down to the level of the water put the tube on one side and examine it from time to time. There are no signs of rust on the nib, so that evidently water alone will not cause iron or steel to rust.



Fig. 10.—Pen-nib in water.

Next let us consider dry air and rust. When is it that steel and iron articles rust? When do they not rust? It will be known to all that it is in damp weather that rust is formed. When the air is dry, as in summer, steel or iron things may be left out in the open without the least fear of rusting. We have thus learned that neither water nor dry air by itself is able to cause iron to rust.

2. Moisten a bright steel nib with water, and expose it to the air for a day. Examine it after this time, and if it has not rusted repeat the experiment.

Eventually it will become coated with rust, so that what air and water cannot do by themselves they readily do when acting in conjunction.

**Explanation.**—Oxygen is slightly soluble in water; and if iron, which will not combine with oxygen directly, is brought into contact with water having oxygen dissolved in it, the water hands the oxygen on, as it were, and under these conditions iron is attracted by and attracts the oxygen, and they combine to form the well-known red rust, which is thus produced by the combination of iron with oxygen, by the aid of water as an intermediary.

3. **Does Iron gain or lose Weight on rusting?**—Tie up some iron borings or French nails in a muslin bag, and weigh the bag with the contained iron. Fix the bag in the upper part of a bell jar, and place the bell jar in water, as in the accompanying diagram. By means of stamp-paper mark the height to which the water rises. Leave the borings (or nails) thus exposed to the action of moist air for several days, after the expiration of which again weigh both bag and borings, and observe the height to which the water has now risen in the jar. How do you account for the difference in weight? Is it in accordance with expectation? Record your results in your laboratory book, and add a few words in explanation of the difference. Now consider the rising of the water in the jar. The water has crept up just as much as it

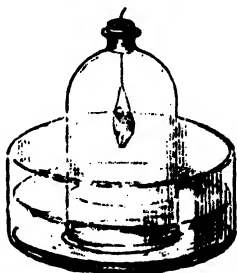


Fig. 20.—Iron rusting in muslin bag.

did when we burned the phosphorus under the bell jar over water. Test the gas that remains with a lighted taper. Which part of air remains, the active or the inactive? Can you understand clearly from this experiment that air is a mixture of active oxygen and inactive nitrogen, four-fifths being nitrogen and one-fifth oxygen, and that it is the active or oxygen fifth that unites with the iron to form iron rust? Just in the same way a stick of phosphorus, if hung in a confined space, will combine with the oxygen, and leave behind the less active nitrogen. The student should perform this experiment, and again test the residual nitrogen with a light and with lime-water. Prove also this same truth concerning the composition of the air by substituting a stick of phosphorus for the iron borings in the above experiment.

4. **Synthesis of Air.**—Place a bell jar over water, and proceed to deprive the air of its oxygen in either of the three ways above referred to—namely, (1) by burning phosphorus; (2) by letting a stick of phosphorus slowly combine with the oxygen; or (3) by exposing iron borings to the air made moist by the water over which it is placed. Or a jar four-fifths full of nitrogen may be prepared by heating ammonium nitrite in a flask, when nitrogen will be given off. Next, from the oxygen apparatus already used, proceed to fill up the jar, and let the two gases thoroughly mix, and then pass on to examine the mixture by burning a taper in it to see if it behaves like air. Any test whatsoever may be applied to it, but no difference between it and atmospheric air can be distinguished. It is fair, then, to assume that by mixing together four measures or volumes of

nitrogen, and one measure or volume of oxygen, air is produced, which affords further proof of the fact that air is a mixture of nitrogen and oxygen, in the proportion of four volumes of nitrogen to one volume of oxygen. It may be pointed out, in passing, that the part played by nitrogen is a most important one. It dilutes the oxygen, and saves us from the dire effects of breathing pure oxygen; an inestimable service, for if we were to breathe pure oxygen the combustion of our tissues would no longer be the slow beneficial combination now in force, but rather an exhausting, energetic combination, the immediate consequence of which would be a much shorter span of life than obtains under the present conditions of slow combustion.

## COMPOSITION OF AIR.

Constituents.	Percentage.	Constituents.	Percentage.
Oxygen .....	20·61	Water vapour.. ..	1·4
Nitrogen .....	77·95	Ammonia .....	Traces.
Carbon dioxide .....	0·04	Ozone.....	Traces.

**Heat and Chemical Combination : Matches.**—Examine an ordinary match and a safety match. Strike them. Why do they burn? Consider first the striking or rubbing. What happens when two bodies are rubbed together, like the match and the box? Everybody knows that heat is produced, and a little reflection will make it clear that the heat really comes from the striker, whose force it is that causes the rubbing. The heat produced by the rubbing or friction should next be considered. What does it do? It makes the match burn. But burning is chemical combination, and we have learned that heat brings about the combination of bodies that otherwise would not combine. Now the burning. Why does this give



us heat? Because it is a rule that when bodies combine to form new bodies heat is given out; and of this we take advantage in making our matches, for we use in making them substances that will readily combine with the oxygen of the air and give out much heat in so combining. What bodies have we already dealt with that burn easily and with great heat? We at once think of **phosphorus** and **sulphur**, and rightly so too, for these are the two chief things used in match-making. In ordinary matches the quick-lighting phosphorus is in the match. In safety matches it is on the box. Write down in your note-book now the reason why safety matches strike only on the box.

## VII. HEAT.

We have been much concerned from time to time with heat, and we shall be again concerned with it in the next section, when we discuss fuels. Let us then ask ourselves what heat is.

### PRACTICAL WORK.

1. Suspend a piece of iron from the beam of a

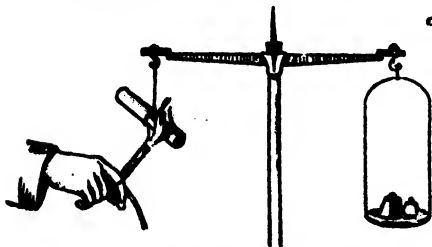


Fig. 21.—Iron does not change in heating.

balance by copper wire, as in the accompanying diagram, and very carefully weigh it. Without removing the iron, proceed to heat it with

a spirit-lamp or Bunsen burner, but do not heat the wire. Does the iron change in weight? If heated cautiously, the answer will be no, and this is always the case when the air does not act chemically upon a heated body; and it proves that heat is not a substance, for it has no weight, and all substance or matter has weight. It is much easier to prove that heat is not substance than it is to prove what it really is; but you may take it for granted that heat is power or energy, and not substance or matter. When bodies combine together, some of the power or energy that was in them is set free, and that is why chemical combination releases heat.

**Fuel.**—Anything that is burned to produce heat may be described as a fuel. There are thus many kinds of fuel—gaseous, liquid, and solid. Of **gaseous fuels**, hydrogen, coal-gas, and carbon monoxide are the chief; of **liquid fuels**, we have petroleum and certain other oils, both mineral and vegetable; whilst **solid fuels**, which are far and away the most important of all, are the various kinds of wood and coal. In burning wood, much of the heat is used up in converting the water contained therein into steam, and in driving off the volatile constituents; hence the conversion of wood into charcoal produces a more valuable fuel, because more concentrated and more powerful in giving heat.

It will be observed that all these are of vegetable origin, and therefore consist of the four elements characteristic of living matter—namely, carbon, oxygen, hydrogen, and nitrogen; and it has already been shown experimentally, and will be shown again by the experiments that are to follow, that when organic

matter is burned in air or oxygen, or allowed to decay, which is really slow combustion, these elements are converted into carbon dioxide, water, and ammonia, the oxygen present assisting in the transformation. We now proceed to examine charcoal, which, as above stated, is the product obtained by heating wood away from air, as in a charcoal-burner's heap, and as in our test-tube experiment with the splinter of wood.

## 2. The Products of the Combustion of Charcoal.—

Place some charcoal in a bulb tube, and connect the

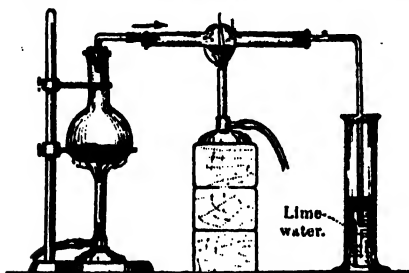


Fig. 22. — Carbon burning in oxygen.

one end with an oxygen apparatus containing chlorate of potash and manganese dioxide, and the other with a vessel containing lime-water. Strongly heat the charcoal

in the bulb tube, and pass a current of oxygen over the heated carbon. The carbon burns brightly, as it did in Experiment 5, p. 113, and at once the lime-water becomes milky, thus proving that carbon dioxide is one of the products of the combustion of carbon.

3. We will now perform this experiment a little more carefully. Take the oxygen apparatus and the bulb tube as before, but now carefully weigh the tube and its contents, and when fitting together, these two parts interpose a number of drying tubes containing calcium chloride between the oxygen tube and the bulb tube. This will ensure that only dry oxygen

passes over the charcoal. Instead of running the gas into the vessel used before, join a set of potash bulbs to the exit end of the bulb tube, the potash bulbs being filled with a strong solution of potash which will absorb the carbon dioxide. Having weighed the potash bulbs, and joined them up as in the diagram, proceed to heat the carbon and send over it a slow, steady stream of oxygen. Note the light produced; and after the experiment has proceeded for some time, take the apparatus apart and reweigh the bulb tube and its contained carbon, and

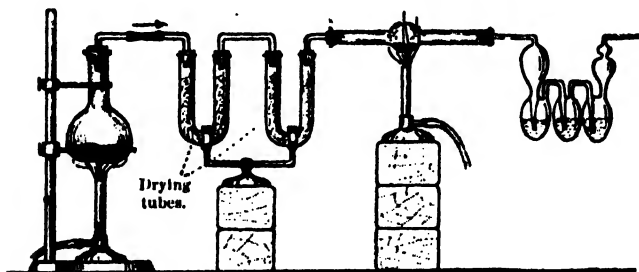


Fig. 23.—Carbon burning in oxygen with potash bulbs.

also the potash bulbs. Record your results, and notice that the gain in the potash bulbs exceeds the loss in the carbon, thus proving that something has been fixed by the carbon in burning. This something, of course, is oxygen. How does this prove that burning is not destruction of matter? Does it prove that burning is chemical combination? If so, how?

4. Taking the apparatus used in either of the above two experiments, but substituting an empty bottle with a stop-cock near the bottom for the oxygen apparatus, proceed to carry out the experiment

with air in place of oxygen, the stream of air being forced over the carbon by the pouring of water into the bottle, which has a thistle funnel passing right down to the bottom for this purpose. The result is pretty much the same as before, except that the combination, as evinced by the less energetic burning, is not so powerful. This falling off is due to the mixture of nitrogen with the oxygen in the air.

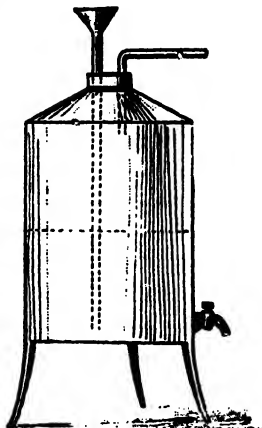


Fig. 24.—Bottle containing air, which is forced out by water.

## VIII. EXPERIMENTS WITH COAL.

### PRACTICAL WORK.

1. Take a long clay tobacco pipe, and place some

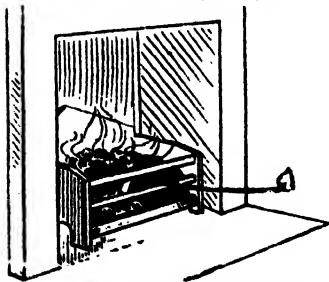


Fig. 25.—Making coal-gas in tobacco-pipe.

of the pipe. What is happening? Is the coal

powdered coal in the bowl, and then place a cap of clay over the bowl, so as to completely close it up. Place the bowl in the fire or in the Bunsen flame. Soon, there appears a stream of smoke from the mouthpiece

burning? No. It is being decomposed by the heat, and the vapours issuing from the mouthpiece represent the volatile portion of the coal. Apply a light to the smoke, and note that it burns with a smoky flame. It is coal-gas in a very impure state that is burning, and our experiment has taught us how coal-gas is made from coal---namely, by heating it away from air; such a method of decomposition being known as **destructive distillation**. What is the difference between heating coal in this way and burning it?

Now pass the impure coal-gas through a U-tube containing water, and again burn it. Observe that it burns much better, the flame being brighter and less smoky. Notice also that the water has become black in consequence of the impurities that have been separated from the gas in its passage through the water. This black substance is **coal-tar**, from which bright colouring substances (aniline dyes), a kind of sugar, and many other important substances, are obtained. This is one of the operations in process at the gas works, where not only has the gas to be produced by the destructive distillation of the coal, but the gas has also to be very carefully purified, or there would be much annoyance and even danger in burning it in our houses.

When the evolution of gas has ceased, take the bowl and examine its contents. These are no longer coal, but, as you will readily guess, **coke**. And now you will understand why the gas manufacturer is a buyer of coal but a seller of coke.

2. Burn a jet of coal-gas from an ordinary gas burner in a large bell jar. You can easily do this by

fitting one end of some rubber tubing on to the burner, and a piece of glass drawn out to a jet on the other end,



Fig. 28. — Water produced by burning coal-gas.

as in the accompanying sketch. Soon some drops of a liquid will collect, and after a time you will have sufficient for the next experiment, which is to boil the liquid that results and to take its boiling point. This will be found to be the boiling point of water, for the simple reason that the liquid is water, which can be further proved by freezing a portion of it in a mixture of ice and salt.

Thus we have proved that one of the products of the combustion of coal-gas is water.

3. Burn the same jet for some time over some lime-water in a bell jar, and shake up the lime-water from time to time, until at last it becomes milky, proving that from the combustion of coal-gas there results also carbon dioxide—the same gas as was given off in yeast fermentation, and when the acid contained in baking powder was made to act upon the carbonate, also contained therein, by the addition of water to the powder. It will be a very useful exercise to compare the gas obtained in these three ways, and to show that it is one and the same.

**Coal: its Production and Chief Varieties.**—Coal is generally found in beds varying in thickness from a few inches to several feet, and extending over wide areas.

Underneath the coal bed lies a layer of clay, called the **underclay**, in which are found roots and rootlets in sufficient quantity to make it impossible to doubt that the coal above them is the remains of ancient vegetation of which they formed the roots. Indeed, in some cases, not only have roots been found, but also the trunks themselves — upright and embedded in the clay wherein, thousands of years ago, they grew and multiplied at so rapid a rate as to prove conclusively that the conditions under which this vegetation was formed must have been exceedingly favourable. It is believed that the vegetation which is now represented by the coal deposits grew in

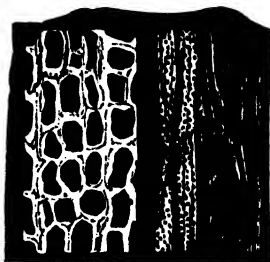


Fig. 27.—Coal under the microscope.



Fig. 28.—Tree trunks in coal bed.

marshy areas, or on the shores of lakes and seas, in very hot climates. Furthermore, the plants themselves were very different from those now to be found



in forests. They consisted of flowerless plants, such as ferns and club-mosses, not, however, the dwarfs that we have with us, but immense specimens rising a hundred feet or more above the ground.

**How Coal was formed.**—The wood, under the conditions of heat and pressure spoken of, lost its gaseous constituents, and became more and more mineral-like in nature, until at last coal resulted. The table given below illustrates the gradual mineralization of wood into coal, and so into graphite or plumbago.

Substance.	Specific Gravity.	Carbon per cent.	Hydrogen per cent.	Oxygen and Nitrogen per cent.
Wood .....	5	50	6	44
Peat .....	75	50	6	38
Lignite .....	1.04	68	5	27
Common (or bituminous coal) .....	1.3	80	5	9
Anthracite .....	1.5	83	3	4
Graphite .....	2.5	100	—	—

**Various Kinds of Coal.**—These, again, illustrate the process of mineralization, as we have, on the one hand, **lignite**, in which the woody nature is conspicuous, and, on the other, **anthracite**, in which there is no trace of vegetable origin, and which does not soil the fingers when it is touched.

**Other Products of the Decomposition of Coal.**—Pitch, asphalt, and petroleum or rock oil are other products of the decomposition of coal. Beneath the famous pitch lake in Trinidad is a coal-field, from the coal of which, by the action of subterranean heat, the pitch is derived.

Petroleum was at one time derived from bituminous shale by distillation; but the discovery of oil springs in Canada in 1861, and later in the deposits

at Baku in Russia, and elsewhere, provided the world with an efficient and cheap illuminant, which, however, should be carefully purified so as to have a high and not a low flash point, as is too commonly the case with cheap oils.

**The Coal Fire.**—From what has preceded, the student ought now to clearly understand that the heat from a fire is due to the combination of the coal with the oxygen of the air, and that the chief products of the combustion are water and carbon dioxide. At the bottom of the grate, where there is abundance of air, the carbon is completely oxidized to the dioxide ( $\text{CO}_2$ ). In the interior of the fire we have red-hot carbon, which breaks some of the carbon dioxide up into carbon monoxide ( $\text{CO}$ ). This

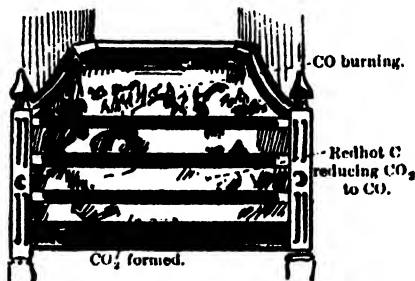


Fig. 29.—The coal fire.

is an inflammable gas, that burns with a blue flame—the familiar blue flame that is so frequently to be seen playing about the top of a clear fire. In this burning the carbon monoxide becomes again carbon dioxide ( $\text{CO}_2$ ).

In making a fire we take care to pack the coal and the wood very loosely together. Why? Both wood and coal must be dry. Why? If the fire does not burn brightly we stir it up. Why? Or we partly cover up the mouth of the chimney. Why? Or we open the door or the window. Why? The fire burns

badly both when the chimney is clogged with soot and when the space below the grate is filled up with ashes. Why? What is soot? and why does it represent great and deplorable waste?

## IX. CANDLE AND GAS FLAMES.

### PRACTICAL WORK.

1. Let us examine the flame of a candle or gas jet. In the first place, insert a piece of glass tubing in the dark portion of the interior of the flame, and apply a light to the end of the tube remote from the flame. The appearance of a flame at this end is a sure proof of the fact that the dark portion of the flame consists of unburnt gas—unburnt for the simple reason that the temperature in this portion of the flame is too low for the combustion of the gas to be carried on; besides, the supply of air is almost nil, which again makes combustion impossible.



Fig. 30.—Unburnt gas in centre of candle flame.

Next take a piece of platinum wire, which is to be placed so that it just touches the flame. Notice that as soon as the wire goes beyond the outer shell or mantle of the flame it goes from a white heat to a red heat. Notice also that this outer shell or mantle gives hardly any light. The remainder of the flame emits a yellowish-white light, and it is evident that to this portion of the flame is due the illuminating effect produced.

**Explanation.**—The outer shell or mantle of great heat and no light worth mentioning is called the **non-luminous zone**. Here the temperature of combustion has been reached, and the air which surrounds it on all sides furnishes an abundant supply of oxygen. Under these conditions it will be readily understood that here is complete combustion, and therefore the maximum of heat, for heat is the direct result of chemical combination, under which, as before stated, combustion is included.

To understand why there is no light, we will return to the luminous zone. Here the gas has been heated to the temperature of combustion, but, owing to an incomplete supply of air, the combustion is incomplete, and solid particles of unburnt carbon resulting from the break-up of some of the gas are in existence. These are raised to a white heat by the combustion of a portion of the gas, and it is to the incandescence of these unburnt particles of carbon that the luminosity of the flame is due. As there are none of these unburnt particles in the outer zone, these considerations account also for the non-luminosity of the outer zone.



Fig. 31.  
Gas flame.

**The Three Zones.**—It will thus be seen that a candle or gas flame may be divided into three zones, which severally are—(1) the dark zone of no combustion, (2) the luminous zone of incomplete combustion, and (3) the non-luminous zone of complete combustion.

The only difference between a candle flame and a gas flame consists in the fact that whereas in the latter case the gas is supplied to the burner ready for combustion, in the former case the solid candle has to be melted and gasified in order to supply the gas required.

**The Bunsen Burner.**—The zone of unburnt gas is unburnt simply for lack of oxygen. If the air is

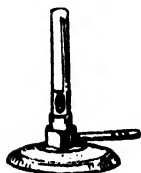


Fig. 32.  
Bunsen burner.

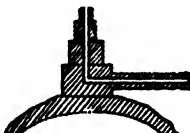


Fig. 33.  
Section of Bunsen burner.

made to mingle with that gas, then the conditions inside the flame are similar to those outside, and complete combustion

tion, with the production of much heat but little light, results. Examine a Bunsen burner. Unscrew the brass tube, and notice the jet inside. Turn on the gas and light the burner, and observe that you have an ordinary luminous flame with the three characteristic zones. Screw on the brass tube, and again turn on the gas, and then light it. The flame now is non-luminous and very hot. Why? Simply because the air passes through the hole at the bottom of the tube and mingles with the gas, so that upon the application of the light there is a sufficiency of oxygen to ensure complete combustion.

**Application.**—Wherever, as in cooking, great heat is wanted and no light, an atmospheric burner, which is simply a more or less modified Bunsen burner, is employed. Examine the burner of a gas stove and see that this is so, and remember that in using a

cooking stove it is of great importance that an adequate supply of air be provided for, so that none of the holes for this purpose should be allowed to be blocked up.

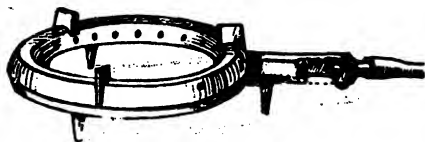


Fig. 34.—Cooking stove.

In an incandescent burner the mantle is composed of a substance something like lime, that may be readily made white hot, or incandescent, without risk of combustion, just like the platinum wire used in Experiment 1, p. 134, and like the lime used in the production of the so-called lime-light. By this means a white light is produced, which is much more suitable for illumination purposes than the yellow light of the ordinary gas burner.

### PRACTICAL WORK.

2. **The Gasalier.**—Take a gasalier to pieces, and observe the peculiar contrivance for lengthening the gas tube without allowing the gas to escape. The outer tube is filled with water, in order to prevent the escape of gas when the tube is lengthened to lower the light.

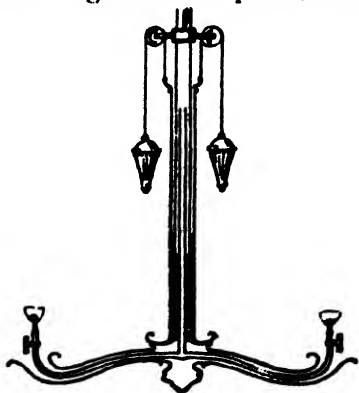


Fig. 35.—Gasalier.

3. **Why Gas Explodes.**—Fill a gas jar with coal-gas by

displacement of water, as formerly explained, and apply a light to it. Notice that it burns quietly, with little or no explosive effect. Next let the gas flow into the jar for a short time, so that the jar shall be filled with a mixture of gas and air; light this, and notice that it explodes. Consider the difference between the two experiments. In the first case, the coal-gas and the oxygen can only combine at the mouth of the jar, as there is no oxygen in the body of the gas itself. In the second case, there is oxygen mixed with the coal-gas throughout its whole volume, so that, immediately upon the application of the light, combination takes place everywhere throughout the mass of the gas, and it is this simultaneous combination with its concomitant expansion that causes the explosion.

**Application to Gas Escapes.**—When there is an escape of gas, every molecule of the gas has near it the oxygen with which it can so easily combine when the temperature is raised sufficiently high. Hence, when the stupid searcher after the place of escape is looking for it with a light in hand, an explosion is bound to take place. The proper thing to do when there is an escape of gas is to throw open

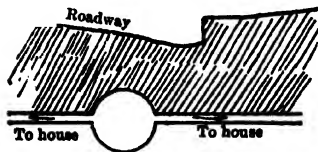


Fig. 38.—Coal-gas main.

windows and doors, so that, by the production of a good draught, the gas collected may be carried off. At the same time the gas should be turned off at

the meter, and, supposing the escape to be a slight one not easily apparent to the eye, a plumber should be

sent for to discover the place of leakage. Should the leak be discovered, then a temporary closure may be brought about by using putty, dough, or white lead.

**The Gas Meter.**—By causing the gas to impinge upon light, easily-turned wheels before its passage into the burner, the number of revolutions of such wheels may be used as a measure of the volume of gas used. This is the principle of the gas meter, which is provided with three sets of dials connected with the wheels, the first of which records hundreds of feet, the second thousands of feet, the third hundreds of thousands of feet. The dials, which are placed on the outside of the meter, since they are connected with the wheels, record the volume of the gas passing through the meter.

**How the Meter Works.**—We must first understand the construction of the meter. Suppose you are looking at the meter from the front face that carries the dials, as in Fig. 37, where a portion of the front face, however, has been removed. Running right across the meter, parallel to the bottom, and at right angles to the front face at which we are looking, is a plate E, which divides the meter into two chambers, an upper and a lower. The upper chamber carries a valve box H, which we shall discuss separately.

The lower chamber is subdivided into two parts by the plate B, which is parallel to the front face at which we are looking.

Attached to this plate are two bellows, C C<sup>1</sup>, one in each of the lower chambers. There are thus four cavities in the lower portion of the meter—namely, the two bellows, C C<sup>1</sup>, and the remaining parts of the





presses upon the bellows, and causes the gas in C to pass out by way of the exhaust hole P into the channel N, and so out of the meter into the service pipes, by O. The diaphragm D, in moving, turns the wire T, to which it is hinged, a quarter of a rotation. This movement of D closes the ingress to A, and opens the passage into A<sup>1</sup>, whereby the bellows C<sup>1</sup> is exhausted, the gas passing through P<sup>1</sup> into O as before, another quarter turn, this time of T<sup>1</sup>, being brought about. The valve, by this movement of D<sup>1</sup>, has caused the opening into C to be freed, and the gas passes from H into C and distends this bellows, the gas being forced out of A into O by the distention, and another quarter turn of T is brought about. Moreover, the action of the valve has by this time opened the passage into C<sup>1</sup>, and the process is again repeated, and another quarter turn registered corresponding to the passage of the gas from A<sup>1</sup> into O. This process of filling A A<sup>1</sup>, C C<sup>1</sup>, with the consequent turns of the valve spindle, is continued, and the turns are by a somewhat complicated arrangement communicated to the clockwork registering gear.

It will be noticed that the movement of the dials is controlled exactly by the movement of the bellows; and the size of these is so arranged that a cubic foot of gas in escaping shall cause the index of the outermost dial to go once round the dial, and so a strict account of the gas passing out of O is kept.

**To read the Meter.**—The figures on the left-hand dial stand for tens of thousands, those in the middle dial for thousands, and those on the right-hand dial for hundreds of cubic feet. To read the index, put down the figure next behind the pointer on each dial,

and add two ciphers. In this way we see that the meter index in Fig. 38 records 16,700 cubic feet of gas. It is to be noted that the index-hand, of the middle dial moves in a direction opposite to that in which the hands of a watch move, whereas the hands of the other two dials move in the same direction as the hands of a watch. If at the next examination the record is 18,300 cubic feet, then by subtraction it

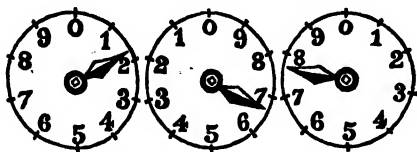


Fig. 38.—Index of gas meter.

is evident that 1,600 cubic feet of gas have passed through the meter in the interval.

The student should read the meter before and after the lighting of the gas as frequently as possible until the system of record is thoroughly grasped.

It will be a very useful experiment to ascertain the amount of gas, the amount of coal, and the amount of oil used in cooking say a leg of mutton, and so to calculate the relative cost of these three different ways of cooking.

## Part III.

### I. WATER.

The most important of all substances to man is water. The shipwrecked mariner and the explorer of the desert both feel most acutely the lack of fresh water. No pang is more excruciating than that of thirst. The reason for this is not difficult to find, for it is well known that two-thirds of the human body consists of water. Or, to put it more clearly, let us take an average man. His weight is 11 stones, or 154 lbs., which is made up as follows:—

- |                       |          |
|-----------------------|----------|
| (a) Dry, solid matter | 45 lbs.  |
| (b) Water             | 109 lbs. |

Again, let us take the blood. Here of water we have more than three-quarters, for of every hundred parts by weight of blood, no less than seventy-nine are water.

Now this water is constantly being drawn upon as long as we are alive. Perspiration by way of the skin, urine by way of the kidneys, and moisture by way of the lungs constantly make the percentage of water in the body less, and therefore the need for taking in

water is one that must be satisfied. Thus water is a food, and of all foods it is the most important.

### PRACTICAL WORK.

1. **The Composition of Water.**—Place some water in a bowl, and in this invert a gas jar also filled with water. Get some thin copper wire, and wind it round a piece of sodium not larger than a good-sized pea. Make the coils of the wire too fine for the piece of sodium to escape from. Throw the sodium into the water, and place the jar over the sodium, keeping the

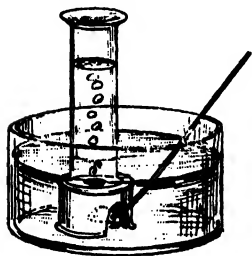


Fig. 1.  
Preparation of hydrogen  
by means of sodium.

mouth of the jar under the water, but not on the bottom of the bowl. Bubbles of a gas pass up the jar and displace the water. It will be noticed that this gas is invisible, and of course it is insoluble in water, or it would not pass up the jar. Take out the jar, and, holding the jar mouth downwards, place a lighted

taper into it. The taper is extinguished, but the gas burns with a very pale blue flame, so faint that it is usually described as non-luminous. The gas that behaves in this manner is **hydrogen**, and our experiment has taught us that water contains hydrogen.

We will spend a little time in examining this constituent of water.

2. **A Better Way of preparing Hydrogen.**—Fit up a flask with a well-fitting cork in which are bored two holes, the one to carry a thistle funnel, the

other a leading tube, as in the diagram. By means of rubber tubing connect the leading tube with glass

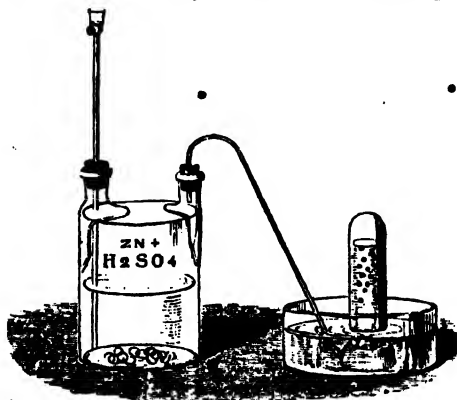


Fig. 2.—Preparation of hydrogen from sulphuric acid by means of zinc.

tubing that passes under a beehive shelf in a trough of water. Before corking put some granulated zinc in the flask. Fill a gas jar with water, and place it inverted on the beehive shelf. Finally, pour

some spirit of salt (hydrochloric acid) down the thistle funnel, and at once a gas will be evolved and will displace the water in the gas jar. Upon examination the gas will be found to behave exactly as that obtained from water by means of sodium. It is hydrogen, which has been obtained from the acid by the agency of the metal zinc. The same result would follow if we were to substitute other metals for zinc, and other acids for hydrochloric, for it

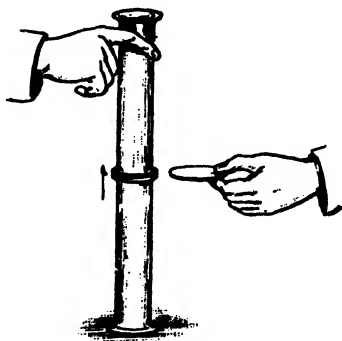


Fig. 3.—Lightness of hydrogen.

is a general law that **metals displace hydrogen from acids.**

Take a jar filled with hydrogen, and on the top of it place another gas jar. After a time apply a light



Fig. 4.—Filling a balloon with hydrogen gas.

to each of the jars. It will be found that the lower jar in which the hydrogen was collected is now empty of hydrogen, and that the gas has passed into the upper jar. Why? Simply because it is lighter than air. A very pretty way of demonstrating the extreme lightness of

hydrogen is to charge a balloon made of very light material with the gas (a collodion balloon is best), and to notice how speedily it rises through the air.

**3. Hydrogen is the lightest body known,** and is therefore well suited for filling balloons. In place of the bent tube that passes under the beehive shelf, put a piece of tube narrowed off to form a jet. Let the evolution of hydrogen proceed for some time, until a portion of the hydrogen collected in a test tube burns quietly. This will occur when the apparatus is free from air; for a mixture of air and hydrogen explodes when ignited, whereas pure hydrogen burns quietly. Now ignite the jet of hydrogen, and let it impinge upon a flask filled with cold water. Drops of a liquid will be

seen to fall and to collect in a beaker placed below for their reception. Sufficient of this liquid should

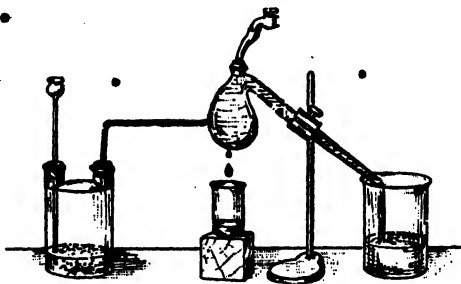


Fig. 5.—Water produced by burning hydrogen in air.

by this means be collected for an examination of the boiling and freezing points to be possible. When this has been done, it will be found that the liquid boils at  $100^{\circ}\text{C}$ . ( $212^{\circ}\text{F}$ .), and that it freezes at  $0^{\circ}\text{C}$ . ( $32^{\circ}\text{F}$ .)—that is, it has the same boiling point and freezing point as water. It is water, and how has it been formed? By the burning of hydrogen in air. But burning is chemical combination, and the active part of air—namely, oxygen—is the body with which the hydrogen combines, so that we may infer from

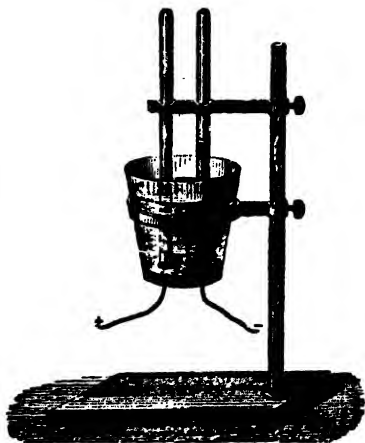


Fig. 6.—Decomposition of water by means of electricity.



this simple experiment that water is a compound of oxygen and hydrogen. There are other more accurate ways of proving this. Thus, if a current of electricity is passed through water, the water is decomposed into oxygen and hydrogen, there being twice the bulk of the latter than of the former; and again, if we pass a stream of hydrogen over heated copper oxide, the hydrogen will combine with the oxygen of the oxide, and water will be formed. This may be collected and

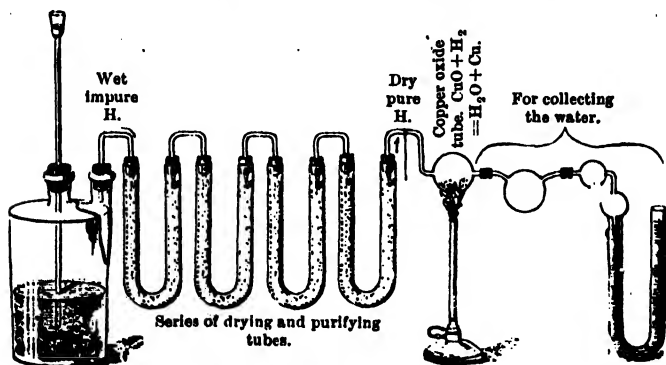


Fig. 7.—Dumas's apparatus for showing that water consists of eight parts by weight of oxygen, and one part by weight of hydrogen.

weighed, and the copper oxide may be weighed before and after the experiment. By these weighings we know (1) the weight of water produced, and (2) the weight of oxygen used. Subtracting the weight of oxygen from the weight of water, we arrive at the weight of hydrogen with which the oxygen combined, and so we are able to state that the oxygen in water is eight times as heavy as the hydrogen. But these experiments we need not perform now, as it will suffice for our purpose that the student should clearly

understand that water is a compound of oxygen and hydrogen, as is shown by the fact that when hydrogen burns in oxygen (or air) water is formed.

4. **The Oxy-hydrogen Flame.**—Take a small lump of lime, and hold it by means of crucible tongs in the flame of burning hydrogen produced as in the last experiment. Notice that the lime becomes white hot or **incandescent**, and, considering the amount heated, gives a very brilliant light. This is the principle of the **limelight**, to produce which a disc of lime is placed in the flame of burning hydrogen, which by heating

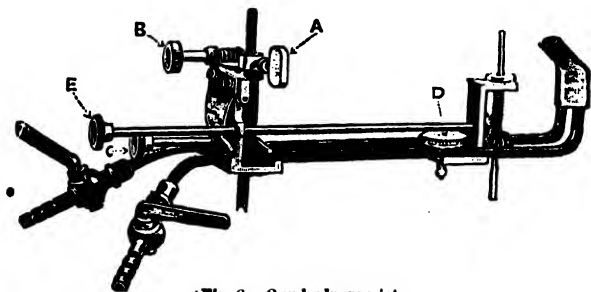


Fig. 8.—Oxy-hydrogen jet.

the lime to incandescence, causes it to emit the brilliant light called on this account the **limelight**.

5. Using the same jet as in Experiment 3, p. 146, send a stream of coal-gas through it and burn this, the flame impinging on the cool flask as before. What do we observe? That drops of water fall as before, the only difference being that the flask becomes coated with a layer of soot or carbon. We may believe then that coal-gas contains hydrogen together with carbon, and that when it burns water is produced. Coal-gas indeed is a mixture of hydrocarbons—that is,

bodies composed of hydrogen and carbon only. If we add to this that the amount of hydrogen present is very great, we may define coal-gas as a mixture of hydrocarbon very rich in hydrogen, from which it will be easy to understand that when coal-gas burns, water and carbon dioxide are formed, as has already been proved.

**Water in Organic Bodies.**—We cannot very well weigh our own bodies, and then desiccate them—that is, drive out the water contained therein—and then, by weighing again, ascertain the proportion of water to solid matter. We can, however, do this with respect to other organic or living matter, and it will be a very interesting and a very instructive exercise to do so.

6. Place a weighed portion of any of the following substances in a drying oven—a water oven will be

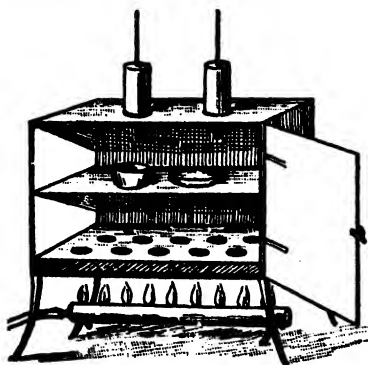


Fig. 9.—Drying oven.

best—and apply heat, taking care, however, that the temperature does not get sufficiently high to prevent the moisture from being driven from the interior of the substance. Weigh again and again, recording the weights thus ascertained in

your laboratory book. It will be found that the same lesson is to be learned in each case, the only difference being in the amount of water present. The following table gives the substances to be experimented upon,

together with the amount of water present in one hundred parts by weight of the substance :—

Substance.	Percentage of Water.	Substance.	Percentage of Water.
Lettuce .....	96	Lean Meat .....	72
Cabbage .....	92	Bread .....	40
Apples .....	83	Cheese .....	34
Fish .....	78	Rice .....	15
Potatoes .....	75	Butter .....	10

**The Amount of Water required by each Individual.—**

For each adult two quarts or five lbs. of water, in some form or other, are the daily requirement. Much of this, of course, is taken as part of the solid food, as a glance at the above table will show, so that the amount of water actually consumed, as water, may be very inconsiderable.

**What the Water Companies allow.**—Taking into consideration the many uses to which water is put—drinking, cooking, washing, and manufactures, it is calculated that a daily supply averaging twenty gallons per unit of population is sufficient for all needs, and it is upon this assumption that the water companies go in fixing what should be the water supply of a place.

**Where Water comes from.**—The original source of all water is, of course, the seas and oceans that take up more than three-fourths of the earth's surface. Every school-girl is familiar with the fact that water passes up from the earth to the sky as **water-vapour** and becomes clouds or water-dust, which, being condensed by cold again, falls to the earth as rain. This circulation is continuous, so that the water of the rivers is touring from earth to sea, from sea to cloud, and from cloud to earth and sea again. This is illustrated in the still, with its double process of

evaporation and condensation, and the student will do well to again perform the experiment on page 70.

Having now made it clear that water is a compound of hydrogen and oxygen, we will proceed with our examination.

7. Fit up a can with a leading tube, and fill it with water, letting the leading tube dip under a beehive shelf, upon which is placed an inverted gas jar filled with water, as in the collection of hydrogen.

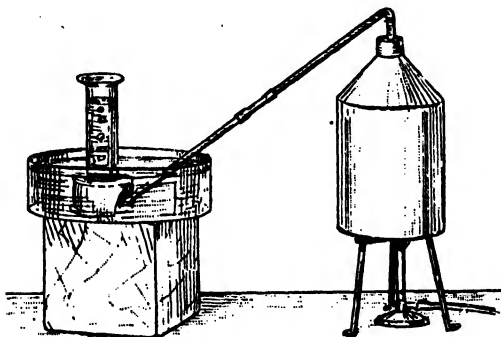


Fig. 10.—Air in water.

Proceed to heat the flask, and observe that bubbles of gas collect in the jar. Where do they come from, and what are they? They are bubbles of air, and they have been driven from the water by heat. It is most essential that there should be air in the water we drink, for otherwise it would taste flat and insipid, as may be at once proved by drinking some of the water from which the air has been expelled.

On board ship, where now most of the water used is obtained by the distillation of sea-water, the water

thus obtained is aerated—that is, charged with air—by running through the perforated bottom of one barrel into another, during which time it takes up the small amount of air that is necessary to make it taste pleasant.

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## II. HARD AND SOFT WATERS.

### PRACTICAL WORK.

1. Fill an evaporating dish with ordinary tap water, and proceed to evaporate the water to dryness by heating it over a Bunsen burner or a spirit-lamp. Observe the sediment that settles at the bottom. This represents the solid matter that was dissolved in the clear, sparkling water.

2. Prepare some lime-water by slaking a piece of quicklime in water, and then pouring off the clear liquid after the lime has settled. Fit up the apparatus shown in Fig. 11 for the preparation of carbon dioxide, and proceed to pass a stream of carbon dioxide into the clear lime-water. Observe that this latter becomes milky or turbid. This is because chalk has been formed by the union of the lime and the carbon dioxide, and chalk is insoluble in water.

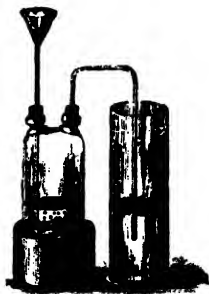


Fig. 11.—Passing carbon dioxide into lime-water.

Let the stream of carbon dioxide continue to flow into the turbid liquid. This after a time becomes quite clear, for although water by itself cannot dissolve

chalk or limestone (they are both the same substance chemically), water charged with carbon dioxide can do this readily.

**Application to Ordinary Water.**—Water in its passage from the clouds to the earth as rain dissolves the carbon dioxide that is always present in the atmosphere. This water, when it falls upon the earth, is at once absorbed by the soil, which accounts for the rapid disappearance of rain puddles after the rain has ceased. If the water in its underground passage comes into contact with limestone or chalk rocks, as is very commonly the case, for chalk and limestone are very widely distributed, the water charged with carbon dioxide can do what we saw it do in the latter part of our experiment—namely, dissolve some of the chalk or limestone, and keep it in solution so long as it is charged with carbon dioxide. Hence when the clear, sparkling water later on bubbles out on a hillside as a spring, it contains quantities of chalk in solution.

3. Take now some of the clarified lime-water and heat it in a beaker. Soon the water again becomes cloudy, and the chalk is no longer held in solution. Why is this? The water upon being boiled yields up its carbon dioxide, just as, in Experiment 7, p. 152, it gave up its air, it being the rule that liquids when heated cannot hold gases in solution to as great an extent as they can when cold. On losing its carbon dioxide, the water also loses its power of dissolving chalk and limestone, which bodies are therefore precipitated from solution, and hence the murkiness of the water.

**The Fur of the Kettle and the Scale of the Boiler.**—

This little experiment will enable us to clearly understand the formation of the layer of earthy substance upon the inside of kettles and boilers. Limestone is the commonest substance found in solution in spring and tap waters; and, as already stated, this substance is only held in solution by the water and the carbon dioxide dissolved therein. Upon boiling, the carbon dioxide is driven out, and the limestone is deposited as scale or fur.

4. **Examination of Kettle Fur.**—Place a portion of the fur of a kettle in a flask fitted with a thistle funnel and leading tube which passes into a vessel containing lime-water. Upon pouring some spirit of salt down the funnel, effervescence at once begins, proving that a gas is given off, and the lime-water soon turns milky, showing that the gas is carbon dioxide. We see then that the fur inside the kettle behaves exactly as do limestone and chalk; and no wonder, seeing that it is the same substance.

5. Take two specimens, the one of rain and the other of spring water. Which is the purer of the two? With which can you most easily wash? Why is this? It is because soap produces a lather much more quickly and much more readily in rain water than in spring water. To prove this, make a soap solution by taking 10 grams of Castile soap and dissolving it in 400 c.c. of methylated spirit, adding after solution 100 c.c. of distilled water. Now pour this soap solution into each of the two specimens of water until a permanent lather results, measure the quantity used in each case, and record this in your note-book.

The reason why more solution is required for the



spring water than for the rain water is that the soap combines with the matter held in solution, and the result is the production of a new body which is not soap, so that, as far as washing is concerned, this amount of soap is lost. As there is practically no dissolved matter in rain water, the whole of the soap used is available for cleansing purposes.

**Application to the Laundry.**—It is very apparent that if two women proceed to wash, the one with spring and the other with rain water, the former will need much more soap than the latter, whose washing will therefore be done at less cost than is incurred by the former; hence the necessity for some cheap way of getting rid of the dissolved matter.

**Hard and Soft Waters.**—Spring water is called **hard** water, rain water is called **soft**, for the simple reason that the former feels much harsher or harder when an attempt is made to wash in it than the former. To soften water, then, is to get rid of the dissolved matter. We have seen that this can be done (1) by **boiling**; (2) by the **addition of soap**. We will now proceed to show a third way, which is much in use in the south of England, where chalk is always contained in the water.

6. Take a specimen of water containing chalk dissolved in it. If the water of your district does not contain chalk or limestone, then proceed to make some artificial hard water, as in Experiment 2, p. 153. Divide this into two parts, and proceed at once to test the hardness of one part by means of soap solution. To the other half add some quicklime, and after the suspended matter has settled down, pour off the clear liquid, and again test with soap solution. \* Why is

less required than at first? Simply because the lime has softened the water by combining with some of the carbon dioxide. We may then give three ways of softening hard water—(1) boiling, (2) soaping, (3) liming, the second being, of course, the most expensive.

**Other dissolved Matters besides Chalk and Limestone.**

—So far we have spoken as though the only matter dissolved in hard water was chalk or limestone. This, as might be expected, is very far from being the case. Water is often described as the **universal solvent**, so many are the substances dissolved by it. But the softening of water containing matters other than chalk in solution is generally brought about by the agency of soap, boiling and the addition of lime having no softening effect. In consequence of this, we speak of **two kinds of hardness**—(1) **temporary**, due to chalk or limestone, which can be remedied by boiling or by the addition of lime; (2) **permanent**, due to other substances, of which calcium sulphate is the most common, which cannot be softened either by boiling or by the addition of lime. For drinking purposes the presence of a small quantity of these mineral matters does not do any harm; indeed, sometimes the effect, in place of being harmful, is **positively beneficial**, as in the case of the **medicinal springs**. But there is one kind of impurity the seriousness of which cannot well be overrated, and that is **sewage matter**. Water polluted by sewage may be pleasing to the eye and in no way offensive to the smell. Cholera and typhoid fever have been frequently traced to contaminated water; and the first care in water provision is to see that it is **absolutely free from organic impurities**. This can only

be done by extreme care in the selection of a supply, so as to ensure a water that has run no chances of pollution. Thus, in the case of London, the water companies taking water from the Thames go well above the metropolis and its environs. For instance, the Lambeth Company gets its supply from Thames Ditton.

**Filters and Filtration.**—On a large scale water is filtered through beds of sand or gravel, the water being poured on the bed and allowed to percolate slowly through the gravel or sand, whereby a double purpose is served, particles held in suspension being separated in the percolation at the same time that the oxygen in the pores of the filtering sand renders harmless any organic matter that may be in the water, by oxidizing it. Care is taken to use the beds only

for a time, so that the air used up may be replaced, to the end that the oxidation may not be stopped.

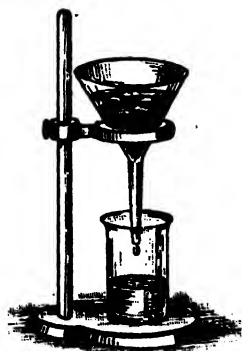


Fig. 12.—Filtration through charcoal.

7. To illustrate filtration, put some animal charcoal in a filter funnel, and pour on it some water containing matter in suspension. The suspended matter does not find its way through. To illustrate purification by filtration, pass a solution of coarse, brown sugar or

treacle through the charcoal, and note that it is decolorized.

**Domestic Filters.**—These are made of animal char-

coal, of silicated carbon, or of spongy iron. Care must be taken to give the filter a rest, as it were, in order that a supply of pure air may be always present in the pores of the filter. If animal charcoal is used, it must not be of inferior quality.

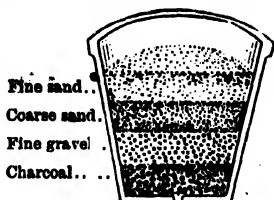


Fig. 13.—Flower-pot filter.

It is always advisable to boil water whenever there is the least suspicion of pollution. To aerate boiled water it is only necessary to pour it several times from one vessel to another.

### III. ACIDS AND ALKALIS.

#### PRACTICAL WORK.

1. **Acids.**—Examine specimens of vinegar, spirit of salt, vitriol, aqua fortis (nitric acid), tartaric acid, citric acid, and oxalic acid. Taste them, making a very dilute solution of the second, third, and fourth, in order to do this safely. They are all sour. Take a few strips of blue litmus paper or some solution of blue litmus, and try the effect of all the acids upon the litmus. The blue becomes red. Compare with this the action of vinegar upon the blue pickled cabbage. Test the three first with zinc, iron, or magnesium. Effervescence takes place, proving that a gas is evolved, which, if the thumb is placed over the test-tube until a sufficient quantity has collected, may be made to burn just as hydrogen did. It is

hydrogen, and this property of yielding hydrogen when acted upon by metals is characteristic of all acids, although it is not always so easy of demonstration as in the cases we have chosen.

2. Procure some solution of ammonia or of caustic potash or soda, and having taken a solution of one or more of the acids, proceed to pour into it a solution of blue litmus, which immediately is coloured red. Now carefully pour into the acid some of the ammonia or caustic alkali solution until the red just begins to turn blue. The acid has been neutralized, and this also is characteristic of all acids that they may be neutralized by alkalis. We have already learned that acids decompose carbonates, and set free carbon dioxide, so that we may now say that acids are sour; that they all contain hydrogen, which may be displaced by a metal; that they are neutralized—that is, made to lose their acidic properties—by alkalis; and that they decompose carbonates. It is to be observed that amongst our specimens are gaseous acids, such as spirit of salt, or hydrochloric acid, liquid acids, as vitriol, and solid acids, as tartaric acid; so that the state or condition—that is, whether the body is a gas, a liquid, or a solid—has nothing whatever to do with the acidic properties.

**Acids are corrosive.**—Most of the acids are corrosive in their action. This is especially the case with vitriol, aqua fortis, and spirit of salt. Vitriol burns holes in clothes, destroys the tissues of the body, and dissolves most metals. This is true, but in lesser degree, of the other two. Some of the acids, such as oxalic acid and prussic acid, are deadly poisons; and as many of them are exceedingly harm-

ful, they should be kept well out of the way of children.

3. **Alkalis.**—Examine solutions of ammonia, caustic potash, caustic soda, and lime. Feel them, and note that they have a **soapy feel**. This is due to the fact that they dissolve the skin. Make a dilute solution of the first three and a strong solution of the last, and wet the tip of the tongue with them. Note the soapy taste. Test them in the same way as you did the acids, only using this time **red** litmus in place of blue. Observe that the **red** becomes **blue**.

4. Take a portion of each of the above solutions in a test tube, and pour in some red litmus solution; then drop by drop pour in some acid solution until the blue is just turning red, when all the alkaline properties will have disappeared, and a new body, neither acidic nor alkaline, will result, as may be proved by evaporating the solution to dryness, and testing with blue and red litmus.

5. Place some olive oil or other liquid fat in water in a test tube and boil. Notice that the globules of fat are not broken up. They simply float about like the soap bubbles in air. Now pour in some of the solution of alkali, and observe that the fats are broken up and that a milky mixture results.

This experiment illustrates the great use of the alkalis to the housewife, since by their aid grease stains may be removed from fabrics. Washing soda, as we shall see later, although not a true alkali, owes its usefulness to the fact that it behaves as a powerful alkali. Thus it would break up the fat globules in the same way as was done by ammonia and the caustic alkalis.

**Acid and Alkali Poisoning.**—As acids and alkalis mutually neutralize each other, it is plain that one can act as an antidote to the other. For instance, suppose a child swallows some acid, say oxalic, the first thing to do is to immediately dose the patient with a solution of ammonia, as the surest means of averting disaster. In like manner, if a liniment or embrocation is inadvertently swallowed by a meddling child or by anybody, recourse should immediately be had to the vinegar bottle, a dose of vinegar being capable of neutralizing some, at least, of the ill effects of the liniment or embrocation, both of which contain much ammonia.

We have proved by our experiments that alkalis have a **soapy taste** and a **soapy feel**; that they turn **red litmus blue**; that they **neutralize acids**; and that they have **very useful cleansing properties**, especially with respect to grease stains.

**6. Preparation of Caustic Soda.**—Place in a large beaker a saturated solution of washing-soda, and then pour in some milk of lime—that is, water in which quicklime has been stirred—until the whole has the consistency of cream. Boil the mixture for some time, and then let it settle. Pour off the liquid that floats on the top of the settled sediment, and then evaporate the sediment to dryness. Compare the product with the caustic soda used in former experiments, and apply to it all the tests that should be applied to a suspected alkali. Show that it is not a soda carbonate by pouring acid on a portion of it, and observing that, although it effervesces because of the admixture of a certain portion of carbonate with it, it does not dissolve, as may be shown to be the case

with soda carbonate. The new body thus produced by the action of washing-soda on lime is **caustic soda**, one of the most powerful of the alkalis. Its name caustic is due to the fact that if it is applied to the skin it will burn it, so powerful is its action. It is used in the manufacture of soap, glass, and many other substances.

#### IV. SOAP AND SODA.

##### PRACTICAL WORK.

1. **Soap.**—Melt some dripping in a beaker, and add about four times its volume of water, and boil. Add about twice the volume of strong caustic soda solution, and boil steadily, with constant stirring. Soap is hereby made, but of course it is dissolved in the water. To get it out of solution, add about half the volume of strong brine, whereupon the soap will be thrown out of solution, and the impure specimen of hard or yellow soap may be separated by filtration and examined. If caustic potash were used, **soft** soap would be the product; and if the alkali lime were employed, a peculiar soap, insoluble, and called lime soap, would result.

We see from this experiment that soap is the body that results from the action of an alkali upon a fat or oil.

The fact that soaps are made by the action of alkalis is brought very unpleasantly home to us if we wash ourselves with common soap, such as the cheap, mottled soap. The free alkali has not been removed, and the smarting of the face after washing



testifies alike to the presence of the alkali and its burning properties.

**Glycerine.**—In speaking of the formation of soap by the action of caustic alkalis on fats, we must not omit to mention the very useful substance glycerine, which is produced at the same time. This cannot be done very easily by the student at this stage, but she can understand something of the process from the following:—Fats are a compound of glycerine and fatty acids; they are, in fact, the glycerine salts of these acids. When, therefore, the alkali acts upon the fat, as in our previous experiment, the glycerine part of the fat is separated, whilst the alkali combines with the remainder to form soap. Thus from the despised fat we get cleanliness-producing soap and chap-dispersing glycerine.

**2. Use of Soap.**—Repeat Experiment 5, page 161, using soap instead of alkali to break up the fat globules. Observe that the result is exactly the same as before.

**Application to the Skin.**—As we shall see later, the pores of the skin exude an oily substance, which must be removed if its healthy action is to continue. Washing in water will not effect the removal of this oily secretion, as we saw in Experiment 5, p. 161. Soap is therefore needed, and by its aid the oil is broken up, and can readily be removed by the towel when the skin is wiped. The action of soap in cleansing clothes, floors, and utensils is exactly as in the cleansing of the skin, fats being broken up and made easy of removal in each case.

**Soda.**—Of all substances valued by the busy housewife soda takes first rank. Over two hundred

thousand tons of it are produced annually in Great Britain, the value of this annual product being not less than two million pounds sterling.

3. **What Soda is.**—Take a piece of washing-soda and put it in a test tube. Pour upon it some vinegar or other acid, and gently tilt the tube, so that the gas which, as the brisk effervescence shows, is coming off in quantity may flow into another test tube containing lime-water. The lime-water becomes milky, showing that carbon dioxide is the gas evolved. Now take a piece of platinum wire, or, if this is not available, a piece of steel wire, and, after dipping it into a solution of soda, put the wire in the Bunsen flame. Observe the rich golden-yellow flame that results. It is the flame of burning sodium, as may be proved by throwing a pellet of sodium in hot water. Soda then contains the metal sodium and carbon dioxide. It is hence known as **carbonate of sodium**, or **soda carbonate**.

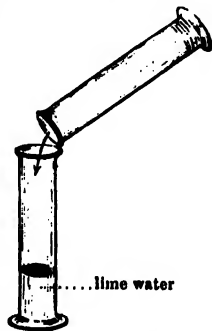


Fig. 14. — Vinegar when poured upon washing-soda decomposes it, and sets free carbon dioxide, as is shown by the milkiness produced in the clear lime-water.

**How Soda is Made.**—We cannot conveniently prepare soda, but we can easily give an idea of the manner in which it is manufactured at the **alkali works**. Sodium is very common indeed, as it is a constituent of the widely-distributed common salt, and it is from salt that the sodium is got for the soda crystals. We have already proved that the widely-occurring limestone contains carbon dioxide, so that the two common substances, salt and lime-

stone, contain all that is necessary to the formation of soda. All that is needed is to get the sodium of the salt and the carbon dioxide of the limestone in combination, and this is what is done at the alkali works.

4. In a test tube put a lump of common salt, and pour upon it some sulphuric acid. Note the cloud that gathers at the mouth of the tube. This is due to the hydrochloric acid that comes off in consequence of the action of the sulphuric acid upon the salt, sodium sulphate and hydrochloric acid, or spirit of salt, being thus formed.

Vitriol and salt yield spirit of salt and sodium sulphate. This sodium sulphate is called salt cake by the alkali worker, and the next step in the process of soda-making consists in acting upon the salt cake with limestone, so that the carbon dioxide of the limestone may be transferred to the sodium, by which limestone and sodium sulphate yield sodium carbonate and calcium sulphate. The crude soda thus formed, called soda ash, is dissolved and purified, and then allowed to crystallize out. How this is done we may now illustrate.

5. Dissolve as much soda as possible in a small beaker or a test tube of boiling water. Let the mixture cool, and crystals of soda will be seen to crop out. This is only a repetition of what we did, in a former experiment.

6. **The Properties of Soda.**—We have already seen that soda is readily soluble in water. Test the solution with red litmus. Boil some fat in it. Show the cleansing action upon grease stains. Why does soda soften water? Why do we put soda into the water

wherein we are boiling vegetables such as cabbage ? Because they boil more readily in soft than in hard water, and soda softens hard water.

**Uses of Soda.**—Besides being used for washing purposes and for softening water, soda is used in the manufacture of glass and caustic soda. In washing it must not be used injudiciously, as it has an ill effect upon delicate fabrics, the fibre of which it weakens. Coloured prints should not be washed in soda, on account of its action upon the dye.

7. Observe the action of soda upon various fabrics, such as muslin, lace, and coloured prints.

**Washing Powders.**—The basis of these is soda, but the much more injurious chloride of lime, or bleaching-powder, is always present; hence, soda is much safer to use than washing-powder, which should be used with caution, as it will have a very ill effect upon delicate fabrics.

**Three different things with similar names :** (1) caustic soda, (2) washing-soda, (3) bread soda (soda bicarbonate).

8. Examine specimens of these three useful substances, all of which we have dealt with in our course of lessons. They are all alkaline, as may be shown by examining them with red litmus.

In a table enumerate the likenesses and differences that exist between these three common substances.

	Water.	Heat.	Litmus.	Acid.	Uses.
Caustic soda.....				•	
Washing-soda.....					
Bread soda.....					

**9. Another Way of Making Washing-soda.**—Dissolve as much salt as you can in a solution of ammonia, or ammonia hydrate, as it is more correctly called. Through this solution pass a current of carbon dioxide, prepared by the action of spirit of salt on chalk. Observe that a white powder begins to separate, and to settle at the bottom of the vessel. This is bread soda, or soda bicarbonate.

**10.** Procure a few grains of this bicarbonate from the shop. (It would be too tedious to prepare it as above.) Heat the powder in a hard glass test tube, and from time to time pour the gas that is disengaged by the heat into another tube containing lime-water. The lime-water becomes milky, showing that carbon dioxide is being evolved. The residue, if dissolved in water, and then allowed to crystallize out, will produce the crystalline soda carbonate used in washing. This is the more common method employed to-day in the manufacture of soda crystals, and, because of the ammonia used in it, is called the ammonia-soda process. It is very desirable that so necessary an aid to cleanliness as soda should be cheaply produced, and Solvay, to whom is due this cheap and simple ammonia-soda process, conferred no small boon upon mankind, for which the housewife will not fail to give him full credit.

**A Twice-told Tale.**—The student will not fail to remember that, in speaking of bread-making and baking-powders, we remarked that carbonate of soda was a constituent thereof, and that the action of the powder in making the bread porous was due to the decomposition of the carbonate by the acid that also formed part of the baking-powder. The soda car-

bonate referred to is not the ordinary washing-soda, but the bicarbonate or bread soda. It is evident from the heating of this substance why we use it in cakes; for when the cake is baked the heat disengages the carbon dioxide, which brings about the desired porosity or sponginess upon which the lightness of the cake depends.

**Manufacture of Bread Soda.**—Bread soda is made by mixing together two parts of washing-soda with three parts of anhydrous soda—that is, washing-soda which has been heated to drive off all the water it contains. Carbon dioxide is passed into the mixture and absorbed by it, and the product is the bread soda or bicarbonate.

**11. To Show that Washing-soda contains Water.**—Place a piece of washing-soda in a test tube and heat it. Observe that it soon dissolves in the water that is thus set free by the heat. Counterpoise a test tube, and within it place another piece of soda, and heat carefully to dryness. Weigh again, and record your experiment in your laboratory book thus—

Weight of soda before drying .....	
" " after " .....	
Loss .....	

As a matter of fact, washing-soda contains more than half its weight of water, differing in this respect from the bicarbonate, which contains no water at all. We give below the chemical formulæ of these two substances, as from these other differences in composition will be apparent:—

• Washing-soda .....	$\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$ .
• Bread soda .....	$\text{NaHCO}_3$ .

It will be seen that there is twice as much sodium (Na) in the former as in the latter compound.

The student should now ask herself why we use soda crystals for washing and not for baking, and why we use bread soda for baking and not for washing.

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## V. STAINS AND THEIR REMOVAL.

**Stains and how to remove Them.**—The good housewife is frequently mortified by seeing clothes, tablecloths, carpets, curtains, etc., rendered unsightly by stains of ink, of grease, of fruit, of paint, or of other foreign substances. To remove these is a very necessary thing to do, and the little chemistry that has already been given will show the way in which she should set to work. The great thing is to find a solvent for the matter that it is required to remove—a solvent, moreover, which will have no injurious effect upon the stained fabric.

### PRACTICAL WORK.

**1. Grease Stains.**—(1.) Take a piece of paper or linen, and stain with candle grease. Place a strip of good blotting-paper on the stain, and apply a hot iron to the part. The heat melts the grease, and the blotting-paper absorbs it. In this way may be removed all grease stains that are easily melted.

(2.) Stain a carpet with grease, and rub the spot with benzine. The spot is removed, because benzine is a solvent for grease, for which reason it is highly valued as a cleansing agent for clothes and woollen material generally.

To cleanse carpets, ox gall is often employed, and wisely so. If a red litmus paper is brought into

contact with the gall, or bile—for that is what it really is—the paper is turned blue; and it will be easy for the student to understand that the cleansing power of gall depends upon its alkalinity, it being a substance similar in this respect to caustic soda and ammonia, being besides safe to use as having no destructive action upon the fabric.

**2. Ink Stains.**—These are very troublesome, and should be dealt with immediately upon their occurrence, by soaking the stained fabric in hot water. Supposing this has not been done, however, proceed as follows:—

Stretch a piece of ink-stained calico over the top of a cup or beaker. Make a strong solution of salt of lemon or salt of sorrel, as potassium oxalate is variously known, and rub the stain until it disappears. Now rinse the fabric well with lukewarm water. Take another piece of calico that has been iron-moulded by ink, and treat it in exactly the same way. A little practice will make you quite adept at removing ink stains, and very useful you will find the accomplishment, for such it is. Instead of salt of sorrel, oxalic acid itself might be used for the purpose of removing the ink stains.

**3. Fruit Stains.**—These are not easy to remove, as fruit juices are true dyes, and will therefore produce a stain difficult of removal. The article stained should be at once transferred to soft water to be well washed, whereby the stain may be removed.

If the stain is an obstinate one, proceed thus: Take a piece of calico stained with damson juice. Stretch it over a vessel, as in the case of the ink stain, and treat it with salt of sorrel as before. Try



the experiment again, this time using salt instead of the salt of sorrel. In each case wash the stain well with boiling water immediately after the application of the salt.

**4. Paint Stains.**—Rub a piece of cloth that has been stained with paint with a mixture of methylated spirit and turpentine. The mixture dissolves out the stain, which may then be rubbed off. Treat the stain next with benzine, in order to cleanse away any grease or other foreign matter.

These experiments should be repeated until the student is well skilled in dealing effectively with these troublesome stains.

**Urgency of Cleanliness in Everything.**—Too much emphasis cannot be placed upon the need for cleanliness in everything connected with the person and the house. Nothing so effectually wards off disease as cleanliness; nothing makes the house more comely and attractive. If utensils are cleaned immediately after use, the cleansing is easier, cheaper, and more effectual than is the case after the utensil has been, even for a short time, placed aside, and although to save a little soap or a little soda may not seem a great thing, "many a mickle makes a muckle" in everything. Again, how unsavoury is the odour from an unclean oven; yet the sluttish housewife would laugh to scorn the suggestion that an oven should be kept just as clean as a saucepan or a frying-pan. This does not alter the fact that dishes cooked in unclean ovens lose much of their savour; hence the necessity for keeping an oven always sweet and clean. So also the stove should be continually attended to in this respect. Blacklead, which beautifies the ugly

rust-disfigured stove, is a sovereign remedy against further rust. It will be of little use to learn the rules and principles of housewifery, if the first and most urgent rule—namely, “Keep all things clean and sweet”—is not obeyed. •

## VI. THE BODY AND FOOD.

**Elementary and Compound Substances.**—It is the business of the chemist to find out whether the substances that are found in the world are made up of one or more kinds of stuff or substance. This he does by breaking up substances into simpler and simpler bodies, until at last he can get nothing more simple. The following experiment will illustrate this:—

### PRACTICAL WORK.

•  
1. Take some red mercuric oxide and heat it in a hard glass test tube. We observe the mirror of mercury that deposits on the side of the tube; and we also, by means of the smouldering chip, show that a gas is given off, which we name oxygen. What do we infer from this experiment? One safe inference is that the red stuff with which we started is not a simple body, but is made up of at least two other bodies simpler than itself, into which it has been resolved by heat. **Mercuric oxide, then, is a compound.**

If we now take mercury or oxygen and experiment with them, we shall find that, do what we may, we can get nothing from them except themselves, for which reason we call these simple undecomposable bodies elements. Thus an element is a simple body

from which nothing has ever been got except itself, such as oxygen, hydrogen, nitrogen, the metals, sulphur, and many other substances.

A compound is a body composed of two or more elements combined together, such as water, starch, sugar, acids, alkalis, etc.

**The Number of the Elements.**—By the most careful and unremitting experiments the chemist has, up to the present time, been able to discover not more than **seventy-five elements**, out of which the whole of the universe, including the animal, the vegetable, and the mineral kingdoms, has been built up; so that the variety of the myriads of different things in sky, earth, and in the waters under the earth, is due, not to the infinite variety of the component elements, but to the amazing number of different combinations that have been made of these seventy-five elements.

**Metals and Non-Metals.**—It is usual to divide the elements into these two classes. By comparing a stick of sulphur with a rod of iron the difference between metal and non-metal can, to some extent at least, be appreciated. Metals have lustre, they conduct heat and electricity, they are malleable, ductile, and tenacious; whereas non-metals are deficient in the qualities above enumerated.

**Three-fourths Metals.**—Of the seventy-five elements, three-fourths (fifty-five) are metals, the rest being non-metals.

**Solid, Liquid, and Gaseous Elements.**—By far the greater majority of the elements are solids. In fact there are but five common gaseous elements—namely,

*oxygen, hydrogen, nitrogen, chlorine, and fluorine; and two liquid elements—namely, the metal mercury and the non-metal bromine—all the rest being solids.*•

These elements unite together to form compounds, and everything in the universe is either an element, a compound, or a mixture of elements or compounds.

**Only a few Elements occur in large Quantity.**—It must not be thought that each of the seventy-five elements is widely distributed; indeed some of them occur in such small quantity as to be chemical curiosities. Further, although certain of the elements, such as oxygen and hydrogen, occur in equal abundance in both living and non-living substances, some of them appear to be characteristic of living and some of non-living or mineral substances. Of this we shall speak later. In the meantime we will endeavour to point out the difference between living and non-living matter, or between organic and inorganic matter, which are the terms usually applied to these two different forms of matter.

**Living or Organic Matter.**—This includes all plants and animals. An examination of a living being shows that it consists of **special parts doing special work**, which work cannot be done by other parts. This is especially the case with respect to the higher living beings, such as trees and men. It is ludicrous to think of a man tasting with his hand or smelling with his eye or seeing with his nose, just as it is to think of a root producing flowers and fruit. These **special parts** are called **organs**, instruments or tools with which a special work can be done; and the

being possessing organs is an organized being, the substance of which the organs are composed being called organic matter.

**A Stone has no special Parts.**—Next consider a stone or a piece of iron. There are no special parts told off to do special work. A stone has no organs; it consists of mineral or inorganic matter.

It is usual to divide organized beings into two classes or kingdoms—the **animal** and the **vegetable**; inorganic matter being all included in a third division known as the **mineral kingdom**.

**Man the most highly-organized Being.**—Of all organized beings, man is the chief; and the next thing that we have to do is to study man, especially with respect to his digestive or food-dissolving parts or organs; after which we shall conclude the course by studying the digestion and preservation of food more fully than has been done in the earlier part of our lessons.

It will be a useful exercise for the student to classify twenty or thirty of the more common substances in her laboratory book under the following heads:—

	Six Compounds.	Six Metals.	Six Non-Metals.
<b>Animal Kingdom.</b>			
<b>Vegetable Kingdom.</b>			
<b>Mineral Kingdom.</b>			

## VII. THE PRESERVATION OF FOOD.

In Experiment 1, p. 94, we showed how, by the growth of a minute organism—the yeast plant—carbon dioxide and alcohol may be produced from a sugary or starchy substance; and how, by the action of another of these germs, vinegar is produced from wine or beer. We called these processes **fermentation**, the yeast and the “mother of vinegar” being known as vegetable ferments.

**Putrefaction.**—We have now to look into the process of putrefaction, that great stumbling-block to the preservation of food. It has been found that putrefaction is simply a process whereby the growth of vegetable ferments produces various unpleasant substances. It has been pointed out that when, as a result of the growth of these vegetable ferments, substances not unpleasant or dangerous result, the process is called fermentation; whereas, if **disease, decay, and death** are connected with the growth of fermenting fungi, we speak of the process as putrefaction.

**Bacteria.**—The minutest fungi are called bacteria. In size they are so small that even after having been magnified four hundred times they are still invisible. What they lack in size, however, they make up for in numbers and in hardness. It has been ascertained that there are not less than eighty to every cubic foot of air. They may be heated up to 100° C., and cooled down to the lowest temperatures, without being killed; although, by this means they are rendered incapable of further growth, until more favourable conditions obtain. When such conditions do arrive, the microbes increase by splitting at an incredible rate, so that

from one millions result. It is only fair to add that of the numerous forms of these lowly fungi but a small minority are deadly to man; some forms appear to be actually beneficial. We cannot further enter into this interesting subject. The student should, however, carefully understand that these microbes are fungi; that they live upon organic or living matter, in which respect they differ from ordinary green plants; that they multiply with extreme rapidity; and that very few substances are able to kill them, amongst which the most common is **carbolic acid**. A substance that has been freed from microbes by carbolic acid or other **antiseptic**, as these microbe-killers are called, is said to be **sterilized**—that is, unfruitful in the growth of microbes.

### PRACTICAL WORK.

1. **Application to Food.**—Expose some milk to the action of the air, or even bottle it up so that the outer air cannot get near it. Observe that after a time it becomes sour. This sour milk now contains a new substance, **lactic acid**, which has resulted from the growth of the spores of a ferment similar to the mother of vinegar ferment. The ferment was present in the milk and is present in the air, and as soon as it grows lactic acid is formed and the milk becomes sour.

Take some more milk and add to it a little sugar, or a little sugar and bicarbonate of soda, or a little boracic acid. Now proceed to boil the milk, and then cork it up tightly. It will remain for any length of time without souring. It is **sterilized**. It will now be evident why it is wise to boil milk and not

to take it raw at any time. There may be other bacteria in it in addition to the lactic acid ferment, and these may develop in the body of the drinker and cause scarlet fever or other dangerous diseases.

**Antiseptics.**—The above experiment illustrates the preservation of food by means of antiseptics—that is, preventives of putrefaction, as the word means. Common salt is the most widely-used antiseptic; **boracic acid** is also much valued. **Creosote** is the antiseptic that is applied to bacon when it is smoked after having been salted. **Salicylic acid** is not so much to be commended as boracic acid; indeed, probably the best antiseptic that can be obtained for ordinary purposes is a mixture of **boracic acid** and **glycerine**, used in the proportion of about 1 to 50 of water, in which it dissolves.

2. Obtain specimens of the various antiseptics referred to above, and test their power in arresting putrefaction by treating meat, etc., with them.

The boiling of milk, water, and other liquids, and the heating of solids, illustrate the fact that substances may be sterilized by having their temperature raised, and if the air is excluded such boiled substances will not putrefy.

**Tinned Meats.**—These are prepared by placing the meat in tins and filling up the tins with gravy, the cover then being placed on, and the whole sealed up *except at one point*. The tin is now heated to a temperature well above that of boiling water, whereby all air is expelled, and all microbes are rendered inactive. The hole is now filled up with solder, and the meat will remain good so long as the air is excluded. To test whether the tin remains air-tight, all



that we now do is to examine the top, which should be bent in or depressed by the pressure of the air. If this is not the case, the meat should not be eaten.

In these days it is well known that meat is exported from New Zealand and Australia in frozen chambers. The cold acts here as the antiseptic agent, which fact is also taken advantage of in the ice-safes that are so much in demand in the summer.

Then, again, who has not heard of dried fish and of dried meat? Drying is a means of preserving food. To sum up, we may enumerate four methods of food preservation:—

(1.) By antiseptics, such as salt, boracic acid, creosote (why not carbolic acid, as it is so excellent an antiseptic?).

(2.) By heating and excluding air, as in tinned meats.

(3.) By refrigerators, as in the case of New Zealand mutton.

(4.) By drying, as in dried fish and the pemmican of Arctic travellers.

With respect to (2), a danger may appropriately be pointed out with regard to tinned fruits. Solder contains lead and tin, and fruits contain acids. Now we have learned from our chemical experiments that metals and acids react, and in the case of tinned fruits the reaction will result in the formation of poisonous salts of lead, etc. The good housewife will, therefore, always buy her preserved fruits in bottles, where there is no solder, and no possibility of the formation of poisonous bodies.

## VIII. DISINFECTANTS.

**Disinfectants.**—The bedroom of a patient suffering from a germ disease (and what diseases are not due to germs?) is a hothouse wherein microbes are multiplying at an alarming rate. Hence the need of disinfection, whereby these germs may be destroyed. It will readily be understood that disinfection cannot be carried on whilst the patient is in the room, since it is much more easy to kill a man than it is to kill a microbe. Hence, in studying disinfection we are only considering what is to be done after the patient has happily recovered, and is in a fit condition to be removed from the room.

1. **As to the Patient.**—He should be thoroughly washed, and should put on a complete change of clothing that has been previously disinfected.

2. **The Room.**—All portable articles, such as beds, mattresses, etc., that have been in close contact with the patient, should be sent to the public disinfector, where they will be effectually dealt with by being subjected to the heat of superheated steam.

The room should now be made **completely air-tight** by pasting strips of paper over window and door chinks, and by filling up the chimney with shavings. About two pounds of sulphur (one pound for each 1,000 cubic feet of space) is now to be powdered and placed on an iron plate, under which there is a bucket, or other vessel, filled with water, and so placed as to catch any burning sulphur that may run off the plate. Strips of paper should be got ready for sealing the door out of which the person who is conducting the disinfection must pass when leaving the room. The

sulphur is now set burning. The operator leaves the room, and immediately stops up keyhole and chinks, and the room is then left to the action of the sulphur for twenty-four hours, after the expiry of which time the door is reopened, windows are thrown up, the chimney is cleared, and everything done to promote a copious supply of fresh air. The paper is completely stripped off the walls and burned, the ceiling is scraped and whitewashed, and every bit of paint, as well as the floor, is scrubbed with carbolic soap; and thus, by the combined agency of disinfectant, fresh air, and cleanliness, the fell enemy is destroyed.

Here the disinfectant used is **sulphur dioxide**, which is the body produced by the burning of the sulphur. **Carbolic acid** and **chlorine** are other less manageable disinfectants.

It should be remembered that the clothes worn by the patient prior to his seizure are probably charged with germs. They also should be carefully disinfected, or the disease may be communicated, notwithstanding the disinfection of the room.

**Disinfectants and Deodorants.**—These two very different things should be carefully distinguished.



Fig. 15.—Absorption of ammonia by charcoal.

### PRACTICAL WORK.

1. Invert a test tube full of mercury in a mercury-trough. Pass up some dry ammonia gas until the test tube is full. Next push a piece of charcoal up the tube, and let the whole remain for some time. Ob-

serve that the mercury gradually creeps up the test tube. Why? Because the ammonia has been absorbed by the charcoal.

In the same way charcoal and other substances will absorb the foul-smelling air of sickrooms, and will destroy the bad odour, but not the microbes. Hence carbon is called a deodorant, because it removes the bad odour. It is not an antiseptic or disinfectant. In the same way camphor, turpentine, burning peat, boiling tar, eucalyptus, Condy's fluid, are deodorants only. Sulphur dioxide and carbolic acid are both deodorants and disinfectants, but in a sickroom they cannot be used in anything like sufficient quantity to disinfect without killing the patient. The same may be said of chloride of lime, which yields a supply of chlorine when acted upon by vinegar or any other weak acid. To make the subject clear, we give below, in tabular form, a list of commonly used disinfectants and deodorants:—

Disinfectants.	Deodorants.
<b>Sulphur Dioxide</b> (From burning sulphur).	<b>Charcoal.</b>
<b>Carbolic Acid</b> (Vaporized by heat).	<b>Camphor.</b>
<b>Chlorine</b> (From bleaching-powder by vinegar).	<b>Turpentine.</b>
<b>Condy's Fluid</b> (When actually in contact, as it is non-volatile).	<b>Burning Peat.</b>
<b>Corrosive Sublimate</b> (Dissolved in water, 1 part in 1,000 parts).	<b>Boiling Tar.</b>
	<b>Eucalyptus.</b>
	<b>Condy's Fluid.</b>
	<b>Sulphur Dioxide.</b>
	<b>Carbolic Acid.</b>

**Infection and Contagion.**—An infectious disease is one that may be communicated from one person or

animal to another, such as influenza, scarlet fever, smallpox, etc. If the transference can only take place by direct contact with the person infected, as in the case of itch, ringworm, and hydrophobia, the disease is said to be contagious. Doctors are agreed that the so-called infectious diseases can be transferred by contact, although the usual way is by conveyance of the germs by the atmosphere, or in some way other than contact. The safest course to adopt with respect to an infectious disease is to isolate completely both the patient and the nurse. This can be most effectively done at a fever hospital, and people are happily becoming every day more and more conscious of the wisdom of sending patients to those centres of skilled nursing and complete isolation. The recovery of the patient is thereby made more certain, and the risk of contagion or infection is reduced to a minimum.

When it is impossible to follow this course, isolation at home should be aimed at. Nobody but the nurse and the doctor should enter the sickroom. Every precaution should be taken to keep food, etc., which has been in the sickroom away from the rest of the household. The room should be made as free of furniture as possible, everything unnecessary being removed therefrom at the outset. A good supply of warm fresh air, constantly renewed, should be insisted upon. All articles used by the sick person should be washed separately in corrosive sublimate solution, and the floor of the room itself should be wiped over with carbolic acid or corrosive sublimate and water.\* In fact, it should be recognized that an enemy is present which can only be successfully fought by using the

best weapons, which are (1) fresh air, (2) cleanliness, (3) disinfectants. Moreover, the utmost vigilance is to be taken, from the outset to convalescence, to prevent the infection from spreading, and this can only be done by complete isolation and utmost vigilance, especially when the patient is about to rejoin the family circle.

The table on p. 186 will give some account of infectious diseases with respect to their symptoms, and also with respect to the possible location of the dreaded germ.

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## IX. GENERAL STRUCTURE OF THE HUMAN BODY.

**Human Physiology.**—The human body is beautifully and wonderfully made! It has been said that if people could see all the actions that are taking place inside their bodies, they would be afraid to move, for fear of disturbing or in some way interfering with the working of the different organs. But there is no need for any such fear. The bodies of all healthy people are so well and strongly made, their parts or organs are so beautifully fashioned and fitted for the tasks allotted to them, that day after day, and year after year they perform their respective functions without disturbance and without failure.

Now, whether our bodies are to be strong and healthy, or whether they are to be weak and diseased, depends chiefly on ourselves. If we breathe foul air, inhabit unhealthy houses, and eat unwholesome food, it is perfectly certain that we shall suffer bodily pain, and that our lives will be shortened. To understand the reasons for this, we must learn something of the

## INFECTIOUS DISEASES.

Disease.	Location of Germ.	Time the Patient should be Isolated.	Time that those exposed to Infection should be Isolated.	Symptoms.	Most Infectious Period.
Measles.	Breath.	Twenty-one days.	Sixteen days.	Running at eyes and nose, with rash on fourth day, first on face.	Early stages.
Scarlet Fever.	Peelings of the skin and the breath.	Eight weeks.	Fourteen days.	Sore throat, hot skin, fever, and vomiting.	Peeling stage.
Chicken-pox.	Breath and pocks.	Four weeks.	Eighteen days.	Slight fever, little lumps.	Convalescence.
Smallpox.	Pustules and breath.	Five weeks.	Eighteen days.	Shivering, vomiting, back pains, fever. Rash and lumps on third day.	Final stages.
Diphtheria.	Breath.	Three weeks.	Twelve days.	Sore throat, weakness, gland swelling.	Throughout.
Mumps.	Breath and saliva.	Twenty-one days.	Twenty-one days.	Pain at the jaw, and parotid gland swelling.	Throughout.
Whooping-cough.	Breath.	Six weeks.	Twenty-one days.	Peculiar "whoop," and vomiting.	Throughout.
Typhus and Typhoid.	Odour from skin. Water, or milk, and stools of patient.	Five weeks.	Four weeks.	Dirty, mottled purple rash fifth day. Headache, bleeding at the nose, purging.	Throughout.

structure or make of the human body, the names and positions of the different organs of which it is composed, and the function or kind of work each organ has to perform. This knowledge is called *Physiology*.

**The Skeleton.**—The bony framework of the body is called the skeleton. It is composed of about two hundred separate bones. At the joints, the bones are joined together by bands of a substance like gristle. In the head there are twenty-two separate bones, only one of which—the lower jaw—is movable; then there is a long row of bones—thirty-three altogether—which form the spine or backbone. To the higher part of the backbone, below the bones of the neck, twelve pairs of ribs are fastened, all of which curve round, so that the upper seven pairs join with the breastbone in front. In each arm and hand there are thirty bones, and exactly the same number in each leg and foot so that we can soon reckon up the entire number of bones—about two hundred—which compose the skeleton.

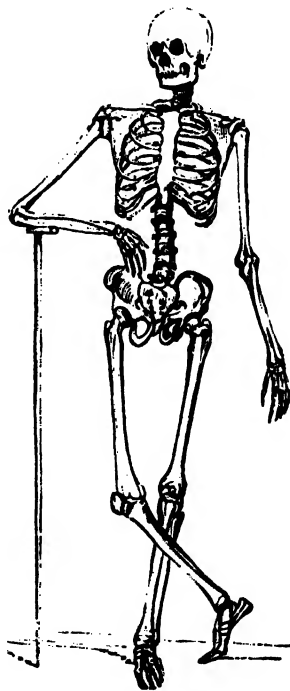


Fig. 16.—Human skeleton.

The use of the skeleton is to form a foundation—a



kind of stiff framework—on which the rest of the body can be built up: the bones of the head enclose and protect the brain; the bones of the trunk perform a similar office for the organs situated in the chest and in the abdomen; while the bones of the limbs impart rigidity to them, and aid us in performing work and in moving about.

**The Muscles.**—The muscles are the great organs of motion. They are composed of fibres laid side by side, forming bundles, which are attached to the bones by

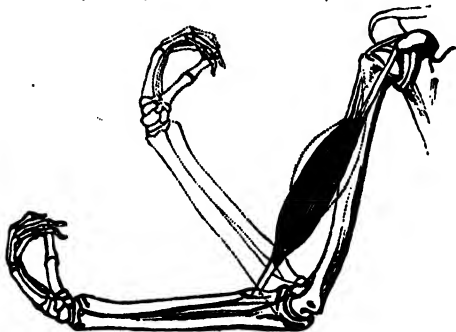


Fig. 17.—Biceps muscle.

tough whitish strings, called sinews or tendons. There is a well-known and powerful muscle, called the biceps muscle, which we can feel forming a lump in the

upper part of the arm when we bend it. This muscle is fastened by two strings or sinews to the shoulder-blade, and by another sinew to a bone in the lower arm, just below the elbow joint. When the biceps muscle contracts, it pulls up the lower arm and the hand. In a similar way the various bones are moved by about four hundred muscles.

**The Organs of Digestion.**—To digest food is to make it fit to enter the blood and nourish the body. The organs of digestion form, or are connected with, a

long tube or pipe, called the alimentary canal. The mouth is the opening into or the commencement of this tube. Next comes the gullet or food-pipe, then the stomach, and lastly the intestines or bowels.

In the mouth the food is broken up by the teeth and mixed with saliva. It is then swallowed, and passing down the gullet it enters the stomach, where a fluid called the gastric juice is mixed with it. Passing out of the stomach, the food enters the small intestines, where two more fluids, the bile (formed by the liver) and the pancreatic juice (formed by the pancreas), are poured upon it. By the action of these four fluids—the saliva, gastric juice, bile, and pancreatic juice—the food is dissolved. Lastly, as the food moves down the small intestines, all the good, nutritious part of it passes through the sides of the intestines into little blood-vessels, along which it is carried, mixed with the blood, to all parts of the body.



Fig. 18. — Organs of digestion :  
a, gullet; b, stomach; c, In-  
testines; d, liver.

**The Organs of Circulation.**—A hot, red fluid is continually moving along certain tubes contained within our bodies. This fluid is the blood, and the tubes through which it runs are called arteries, veins, and capillaries. But what causes the blood to move? A very strong hollow muscle called the heart. The arteries are pipes leading *out* of the heart, the veins are very similar pipes leading *into* the heart, while the

capillaries are extremely small narrow tubes connecting the ends of the arteries with the ends of the veins. Let us suppose the heart to be full of blood. Now the heart *contracts*, driving the

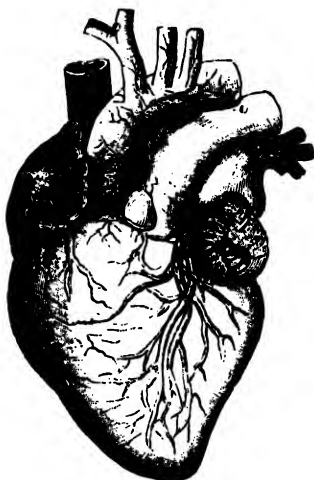


Fig. 19.—Heart.

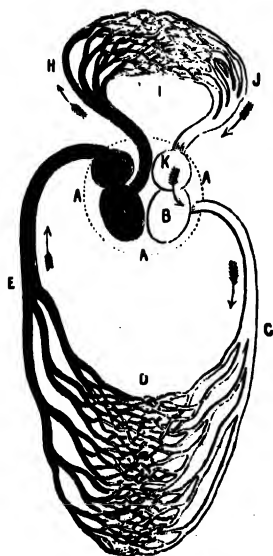


Fig. 20.—Circulation of the blood: A, A, A, the heart; F, G, the two chambers which form the right side of the heart; K, B, two chambers which form the left side of the heart; H, artery conveying blood from right side of the heart to the lungs; J, vein conveying blood from lungs to left side of heart; C, great artery conveying blood to the capillaries; E, great vein bringing blood back to the heart.

blood into the arteries, which carry it to every part of the body. After passing through the capillaries, the blood enters the veins, which pour it into the heart again. In this way the blood is continually *circulating* or flowing round and round within our bodies.

### The Organs of Respiration.

—By respiration we mean the act of breathing. When we take in air, it passes through either the mouth or the nostrils down a tube,

called the windpipe, into the lungs. The lungs are situated in the chest or thorax, one on the right and the other on the left side of the heart. They are light bags, elastic and spongy in texture, and very full of capillaries. The blood is sent from the right side of the heart to the lungs, through a tube called the pulmonary artery. The blood in this artery is dark in colour, and contains much carbon dioxide gas. The blood

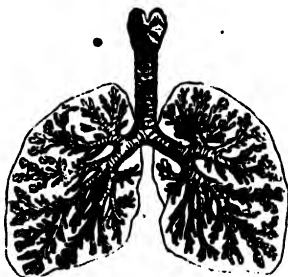


Fig. 21.—Air passages of lungs.

is taken away from the lungs by the four pulmonary veins, which lead it back to the left side of the heart. The blood in these veins is of a scarlet hue; it contains *less* carbon dioxide gas and *more* oxygen than the blood in the pulmonary artery. It is certain, then, that in the lungs the blood *loses* some of its carbon dioxide and *gains* some oxygen. From what source does the blood obtain this oxygen? and what becomes of the carbon dioxide? It is from the air which enters the lungs that the oxygen is taken. Some of the oxygen gas passes through the moist bladdery lining of the lungs *into* the blood; while a nearly equal quantity of carbon dioxide passes from the blood *outwards* into the air, and is expelled from the mouth or nostrils. The air we breathe *out* contains about five per cent. *less* oxygen, and about five per cent. *more* carbon dioxide, than the air we breathe *in*. In this way we take from the air about a pound and a half of oxygen gas every day, and we daily

pour into the air about two pounds weight of the heavy, poisonous gas called carbon dioxide.

**The Organs of Sensation.**—Every one has heard of the five senses: we hope that each reader of this book possesses them all. One sense—namely, the sense of touch—resides in the skin, and is possessed by the entire surface of the human body. But the other four senses are severally restricted to very small portions of the body: thus, the eyes are the organs of sight, the ears are the organs of hearing, the nose is the organ of smell, and the tongue and palate the organ of taste.

**The Nervous System.**—All the organs of the body are connected, by nerves (which look like white threads), either with a great mass of nervous matter, called the brain, or with a long thick nerve, called the spinal cord, which runs down the centre of the backbone. It is by means of the nervous system that our muscles are set in motion when we move. If the nerves going from, say, the tip of a finger to the brain are cut, we can no longer feel anything with that finger. If the brain is diseased, we cannot think properly: the brains of idiots and of mad people are always found to be defective in some respect.



Fig. 22.—The brain and spinal cord with its branches.

**The Organs of Excretion.**—To excrete is to remove and get rid of useless matter. The food we take does not accumulate in the body; if it did, we

should grow heavier day after day, until we became perfect mountains of flesh. Each portion of food does its allotted work. Perhaps it goes to produce heat, perhaps to make muscle, perhaps to enable us to perform work; but sooner or later it becomes used up and converted into waste matter, which must be got rid of, as it would be a serious and even dangerous clog if it were not removed from the body. The chief organs of excretion are the lungs, the skin, and the kidneys.

*The lungs* remove carbonic acid gas, water-vapour, and heat from the blood. Much more of all these is contained in the air we breathe *out* than in the air we breathe *in*.

*The skin* excretes the sweat, a liquid which is chiefly water.

*The kidneys* also excrete water, but in this water there is dissolved a nitrogenous substance, called urea.

### PRACTICAL WORK.

Breathe out through a glass tube dipping into clear lime-water: the latter is rapidly turned milky by the carbonic dioxide in the breath. The presence of water-vapour in expired air can be shown by breathing on any bright, cold piece of metal, as a silver spoon, while the mercury in the thermometer quickly rises to about  $98\frac{1}{2}^{\circ}$  when we breathe on that instrument.

**The Organs of the Body all work together.**—Although we see that the human body is made of many distinct parts or organs, each of which has a special work to do, we must not think that they are independent of

one another. All the organs work together for the good of the whole body.

The muscles move the bones and carry the body about, in obedience to orders from the brain, sent through the nerves; the alimentary canal makes the food fit to nourish every organ; the organs of circulation convey the nutritious part of the food, mixed with the blood, to every part of the body; finally, the organs of excretion remove the used-up matters from the body.

But the organs cannot work properly unless they are in a healthy state. And when any organ is unable to do its work well, our body feels uncomfortable and our health suffers. Moreover, since the organs "all work together," when one of them is diseased all the rest suffer. It is, therefore, of the greatest importance that we should understand the nature and the working of our organs, so that we may be able to treat them properly and not to injure them—as many people who know no better do—by over-exertion, over-eating, drinking to excess of alcoholic beverages, want of exercise and fresh air, etc. Just as every engine-driver ought to understand the construction of his engine, so every one who inhabits a "body" (as we all do) ought to understand its build and mode of working.

**Composition and Weight of the Body.**—How many of the seventy elements help to form our bodies? No fewer than sixteen have been detected in the human body, but of four of these there are only a few grains.

Suppose we take the average weight of a grown-up man to be 11 stones, or 154 lbs., then the body of that man would be composed of such weights of the different elements as are stated in the following table.

**Elements forming the Body:—**

	lbs.
Oxygen (a gas).....	109
Carbon (a solid ; non-metal).....	18
Hydrogen (a gas).....	14
Nitrogen (a gas).....	5
Calcium (a metal).....	4
Phosphorus (solid ; non-metal).....	1½
Sulphur (solid ; non-metal).....	½
Chlorine (a gas).....	½
Sodium.....	½
Potassium.....	½
Magnesium.....	½
Fluorine.....	½
Iron }	A few grains.
Silicon }	
Manganese }	A trace.
Copper }	

**Compounds in the Body.**—Of the sixteen elements named in the list above, oxygen is the only one which we find in the separate state, *as an element*, in the human body; and even of free or separate oxygen there is very little. The elements join together to form *compounds*, such as water, common salt, etc.; the compounds form *tissues*, such as muscle, fat, bone, etc.; of such tissues the *organs* are composed; while the *body* consists of all the organs taken together.

The following are the principal compounds of which the human body is made:—

	lbs.
Water (composed of Hydrogen and Oxygen).....	109
Dry Protein Matters (composed of Carbon, Hydrogen, Oxygen, and Nitrogen).....	29
Dry Fatty Matters (composed of Carbon, Hydrogen, and Oxygen).....	5
Dry Mineral Matter (phosphate of lime, common salt, etc.).....	11
	<hr/> 154

The term *protein* means the same as *albuminoid*, or *nitrogenous organic matter*. The *proteins* in the body include *fibrin* in the muscles, *albumen* in the



muscles and in the blood, and *ossein*, the soft organic matter in bones, which can be obtained from them by long boiling, by which it is converted into *gelatin*. The proteids are composed mainly of the four elements carbon, hydrogen, oxygen, and *nitrogen*, but they also contain a little sulphur and phosphorus.

*Fat* is composed of carbon, hydrogen, and oxygen, but it contains very little of the last-named element. It is found in masses round the kidneys, and there is usually a layer of fat just under the skin.

*The mineral compounds* in the body occur chiefly in the bones, to which they impart hardness and rigidity. Of phosphate of lime (composed of phosphorus, oxygen, and calcium) there are  $8\frac{3}{4}$  lbs.; of carbonate of lime, about one pound; of common salt (composed of sodium and chlorine), about half a pound. Salt is always present in the blood.

*The tissues.* A layer of fat, a piece of bladder or skin, and the stringy part of lean meat, are good examples of tissues: we should call them fatty tissue, connective tissue, and muscular tissue, respectively. Of layer upon layer of such tissues all the organs in the body are composed.

**Table showing the average Weight of the Organs of which the Body is composed:—**

	lbs.
Muscles .....	63
Skeleton .....	24
Skin .....	10½
Fat .....	28
Brain .....	3
Blood .....	912
Organs in the Chest (Heart, Lungs).....	2½
Organs in the Abdomen (Stomach, Intestines, Liver, Pancreas, Spleen, Kidneys).....	11
<b>Total.....</b>	<b>154</b>

**Work done in, and by the Body.**

*Work done in the body.* Suppose we took a man who weighed exactly 154 lbs., and placed him in one pan of a large pair of scales, putting his exact weight into the opposite scale-pan. If the scales were constructed very accurately and carefully, we should soon see the weights begin to move *down*, while the scale-pan with the man in it would rise *up*. This would show that the man was *losing weight*.

How can we account for this? If we made the man breathe out through some lime-water, we should see the liquid turn milky: this shows that the man's body is parting with, or giving off, the heavy gas called carbon dioxide. If we let his breath fall on a piece of looking-glass, we should see the water-vapour which is also going out of the body in his breath. We should also find more water-vapour evaporating into the air from the man's skin. Of course, these things—the carbonic acid gas and the water-vapour—have weight, and since they pass away from the man's body into the air, they cause him to lose weight. It is true that the man takes in oxygen gas from the air, but the weight of this is not so great as the weight of the matters he parts with.

Then the man *loses heat*. This must necessarily be the case, since the heat of the healthy body is always  $98\frac{1}{2}^{\circ}$ , while the heat of the air around us is, on the average, only  $50^{\circ}$ .

The man may not have stirred hand or foot since we put him in the scale, yet the production of heat, water, and carbonic acid gas shows us that work is being done in his body. A dead body would not lose weight, for no work would be going on in it. But

when we examine closely the living body, we can see that work *is being done in it*. The rise and fall of the chest and the beating of the heart show this; the pulse proves that the blood is coursing through the arteries; the warm breath proves that something (the food) is being burned within the body, producing heat.

*Work done by the body.* Now let the man jump out of the scale, and set him to perform some hard work for an hour. On weighing him again, it will be found that he has lost much more weight than he did in the same length of time when sitting still. Men who are engaged in very hard, heavy work, which they have to perform in a hot atmosphere—such as puddling iron—have been known to lose several pounds in weight in two or three hours. When we walk about, our muscles have to do work; they have to transport the weight of the body from one place to another.

**All Work implies Waste.**—We here use the word “waste” in the same sense as when we speak of anything wasting away—that is, being used up.

Let us imagine a steam-engine with its boiler filled with water and its furnace full of coals: the whole weighs, say, 20 tons. Now we light the fire, and the engine begins to work; in an hour or two, however, the motion slackens, and at last it stops.

Now we weigh the engine again: it weighs perhaps 18 tons. What has caused it to lose 2 tons in weight? We look into the boiler: it is empty; the water has all become steam, and has passed out into the air. The coals too have disappeared; they have been changed chiefly into carbon dioxide, which, like the steam, has passed away, while the ashes, which would not burn, have dropped through the furnace bars on to

the ground. Even the engine itself has worn away to a slight extent. The parts made of iron or of brass which rub against one another when the engine is at work need repairing from time to time.

It is the same with every other machine. *All work implies waste.* Work cannot be performed without using up something.

**How the Waste of the Body is repaired.**—The human body may be well compared to a railway engine. Like the engine, the body must be built up and kept in repair; but it is able to do this work for itself, which the engine cannot do. If the engine is to do work, we must supply it with water and with coal. It is the same with the human body, which cannot do work unless it is supplied with food; and just as the engine uses up the coal and the water, so does the body use up the food.

## X.—FUNCTIONS AND CLASSIFICATION OF FOODS.

**Functions of Food.**—By the word “function” we here mean *use* or *work*. How does the food act? What does it do in the body? and how does it do it? These are the questions we have to answer.

Considering the matter broadly, we may say that the functions of food are four in number:—(1) To form or build up the body; (2) to keep the body in repair; (3) to maintain the heat of the body at a temperature of  $98\frac{1}{2}^{\circ}$ ; (4) to produce force, thereby enabling the body to do work.

Now there is no single food—except milk (for young children)—which performs all these four functions satisfactorily. One kind of food, it is found, acts

## 200 FUNCTIONS AND CLASSIFICATION OF FOODS.

chiefly in forming new tissues; another kind is best able to produce heat; and so on. Before we begin to examine the functions of the different foods we take, we shall classify them according to the work they perform.

### Classification of Foods.

#### NITROGENOUS FOODS: Flesh-Formers Mainly.

<i>Name.</i>	<i>Food.</i>	<i>Chemical Composition.</i>
ALBUMEN .....	Eggs, meat, etc.....	Nitrogen combined with carbon, hydrogen, and oxygen, as $C_{72}H_{112}O_{22}N_{18}$ .
CASEIN .....	Cheese.....	
FIBRIN.....	{ Lean meat (muscle), oatmeal, etc.....	
OSSEIN.....	Bones (soft part).....	
LEGUMEN .....	Beans, pease, etc.....	
GLUTEN .....	Wheat-flour.....	

#### CARBONACEOUS FOODS:

##### Heat-Givers and Force-Producers Mainly.

<i>Name.</i>	<i>Food.</i>	<i>Chemical Composition.</i>
AMYLOIDS .....	{ Starch ( $C_6H_{10}O_5$ )..... Sugar ( $C_{12}H_{22}O_{11}$ ).....	{ Carbon combined with hydrogen and oxygen.
FATTY FOODS....	{ Oils ( $C_{57}H_{104}O_6$ )..... Fats ( $C_{57}H_{110}O_6$ ).....	{ Carbon combined with hydrogen and a little oxygen.

#### MINERAL OR INORGANIC FOODS.

<i>Name.</i>	<i>Food.</i>	<i>Chemical Composition.</i>
PHOSPHATE OF LIME .....	{ Vegetables.....	{ Calcium combined with phosphorus and oxygen.
OXYGEN .....	Air.....	Oxygen.
WATER .....	Water, vegetables, etc.	{ Hydrogen combined with oxygen.
COMMON SALT ...	{ Salt and vegetable and animal foods.....	{ Sodium combined with chlorine.

This table shows that we take only one element in the separate state as food—namely, oxygen; all our other foods consist of compounds of two, three, or more elements.

The table shows, too, that we obtain food from all the three kingdoms of nature. Plants live on mineral matter only; they make their tissues out of mineral

substances which they obtain from the soil, the water, and the air. Animals cannot do this. Some animals, called *herbivorous*, live on plants only—the cow, for example; other animals, as the lion, are called *carnivorous*, because they feed upon flesh. Man is both herbivorous and carnivorous, eating both vegetable and animal food, and much mineral matter as well, and is therefore said to be *omnivorous*.

We will now examine the three great classes of foods in turn, commencing with those whose principal office or function it is to form new flesh.

**Nitrogenous Foods.**—These are also known as proteids, albuminoids, and flesh-formers. They are composed of the four elements carbon, hydrogen, oxygen, and nitrogen, and generally contain also a little sulphur and phosphorus. They are very complex substances. In the preceding table, the composition of a single molecule (or smallest possible piece) of any proteid is given, and it is seen to contain two hundred and twenty-four atoms!

The proteids are known by different names—such as albumen, fibrin, casein, and gluten—but there is very little difference indeed in their chemical composition.

TABLE OF PROTEIDS.

Name of Food.	Name of Proteid (or albuminoid) contained.	Weight of Nitrogenous matter in 100 lbs. of the food.
Bread.	Gluten.	8
Oatmeal.	Fibrin.	14
Pulse (beans, case, and lentils).	Legumen.	23
Cow's Milk.	Casein.	4
Eggs.	Albumen.	14
Mutton.	Fibrin and albumen.	12
Beef.	Fibrin and albumen.	10
Poultry.	Fibrin and albumen.	12
Fish.	Fibrin and albumen.	13

**Quantity of Nitrogenous Food required daily.**—It is found that we lose, or part with, about 300 grains of nitrogen daily. This mostly leaves the body in a substance called *urea*, which is found dissolved in the water removed from the blood by the kidneys. To replace this, we ought, of course, to eat 300 grains of nitrogen daily. If nitrogen by itself, uncombined, were of any value as food, we should be able to get an ample supply from the *air*, four-fifths of which consists of the colourless, transparent, incombustible gas called nitrogen; but the only shape in which nitrogen can nourish us is as part of some albuminoid substance. Now 2,000 grains weight of albumen or of casein, for instance, will give us 300 grains of nitrogen. But, again, there is no food which is pure albumen or pure casein, etc.: eggs, for example, only contain 14 parts in 100, or about one-seventh of their weight, of such proteid matter as albumen. To get the 300 grains of nitrogen from eggs, then, we must eat 14,000 grains weight ( $2,000 \times 7$ ), or about eighteen eggs daily. The food which is richest in nitrogenous matter is pulse, which contains about one quarter of its weight of legumen: by eating 8,000 grains of lentils ( $2,000 \times 4$ ), we can obtain the 300 grains of nitrogen to supply our daily wants.

In practice, however, we do not obtain our nitrogen from any one source, but from several—bread, meat, milk, eggs, etc., all of which contain proteid matters.

**How Nitrogenous Food is digested.**—Let us take some nitrogenous substance, such as lean meat, and see how it is digested. In the mouth it is masticated, or broken into small pieces, by the grinding action of the teeth. Here also it is mixed with saliva, which

moistens the meat, and makes it easier to swallow, but has no other effect upon it. Now the meat is swallowed, passing down the gullet into the muscular bag called the stomach. Here it is *dissolved* by an acid fluid, called the *gastric juice*; and much of the liquid so formed is *absorbed* by the sides of the stomach, passing into the blood-vessels there. The remaining portion of the dissolved meat passes out of the stomach, and is similarly absorbed by the blood-vessels in the sides of the small intestines.

**Functions of Nitrogenous Foods.**—The main functions of nitrogenous foods are—(1) to furnish material for the construction of the various organs of the body, especially the muscles; and (2) to maintain them at their proper size and weight when fully grown. A baby boy one month old weighs perhaps 12 lbs.; when he is ten years, he perhaps weighs 70 lbs.; and when thirty years old, as much as 154 lbs. All his organs have been steadily increasing in size and weight; and it is the nitrogenous food which has furnished most of the material of which to build them up. After thirty years of age, the man does not perhaps increase in weight; but his organs are continually working, and therefore continually wasting away, and now the nitrogenous food supplies fresh material to make good the loss. The proteids are carried about in the blood, and where a little bit, perhaps of a muscle, has done its work, and has been removed by the blood, there the fresh fragment of proteid matter steps in and fills its place.

But nitrogenous foods are not only flesh-formers. By the oxidation of the carbon and hydrogen which they contain they also *produce heat, and do work*.



Still their *chief*, and by far their most important, function is to form new flesh.

We have compared the human body to a steam-engine; but it might also be compared to a *building*, such as a school. The school was built up, brick by brick; and so is our body built up by the nitrogenous foods which we eat. The school slowly wears or wastes away; the bricks fall down, the slates tumble off, and the windows break. But it is kept in repair by men who put in new bricks, and new slates, and new panes of glass. And so the body is kept in repair by the blood, which carries with it plastic, or tissue-making, or nitrogenous matter; and wherever a bit of muscle has wasted away, there some of this proteid matter is deposited to take its place. Again, the school is warmed by the combustion of carbon (coal); and so is the body warmed by the combination of the carbon in our food with oxygen inhaled from the air.

## XI.—CHARACTERS AND FUNCTIONS OF CARBONACEOUS FOODS.

**Nature of Carbonaceous Foods.**—In nitrogenous foods the most important element is nitrogen; but the foods whose functions we have now to study contain no nitrogen, being composed of carbon, hydrogen, and oxygen. They are named carbonaceous, because carbon is the most important of these three elements; sometimes they are called non-nitrogenous foods, from the total absence of nitrogen.

**Characters of the Elementary Body Carbon.**—The element called carbon is familiar to us in several

forms. It is a solid, usually of a black colour, forming coke, charcoal, graphite (also known as blacklead and plumbago), lamp-black, bone-black, and soot; when crystallized, carbon forms the precious sparkling gem called the diamond.

**Composition of Carbonaceous Foods.**—Carbon by itself is useless as food. A man might take a quantity of charcoal or of coke into his stomach, but he could not digest it, and it would do him harm instead of being useful to him as food.

When combined, however, with hydrogen and oxygen, carbon forms a variety of foods which are pleasant to the taste, easily digested, and fitted to do a very useful work in the body.

The following table shows the **Principal Carbonaceous or Non-nitrogenous Foods** :—

CARBONACEOUS FOODS.

Fatty.	Sugary or Starchy (Amyloids).
(1) SOLID FATS— Butter, etc.	(1) Saccharine or sugary— Cane sugar, etc.
(2) LIQUID FATS or oils— Olive oil, etc.	(2) Starchy— Sago, tapioca, etc.

All the foods named in the above table are composed of the same three elements—carbon, hydrogen, and oxygen—but they differ widely in the proportions in which these elements are mixed together.

*Fatty foods* contain much carbon and hydrogen, but very little oxygen.

*The amyloids* contain much more oxygen; just enough, in fact, to form water if combined with the hydrogen in them. (See page 207.)

**Functions of Fatty Foods.**—As the work done by the

fatty foods is somewhat different from that done by the amyloids, we shall consider separately the functions of each class. How rapidly and brightly a piece of fat burns when placed on the fire! Oil, too, will burn well, and we use it in lamps to give us light and heat. But what do we mean by *burning*? It is the uniting of the carbon and the hydrogen in the oil or fat with oxygen from the air: the carbon and oxygen unite to form carbon dioxide,  $\text{CO}_2$ ; the hydrogen and the oxygen combine to form water-vapour,  $\text{H}_2\text{O}$ . Now when the fat of our food gets into our blood it burns there, combining with the oxygen which has been taken into the blood through the lungs. It is true that this burning in the blood is a very slow process, and that the heat produced is not very great—only  $98\frac{1}{2}^\circ$ ; but this temperature is maintained for a long time, and the total quantity of heat produced by burning a pound of butter is the same, whether we burn it quickly in a few minutes on a fire, or spend some days in burning it in our blood.

The first and most important function, then, of fat as a food is *to produce and maintain the heat of the body*.

Instead of burning our piece of fat on an ordinary fire, we might have burned it in the furnace of a steam-engine: the heat it produced would then have made the water in the boiler hot, and enabled the engine to do some work. Now this is another function or work which fatty foods perform in our bodies—they enable us *to exert energy and to perform work*. It is found that the harder we work, the more carbon dioxide comes out of our mouth; but this substance cannot be formed except by the combination

of carbon and oxygen, and most of the carbon used in the body for this purpose comes from the fatty foods.

Fatty foods also *help to form fat*. When we eat any kind of fat, as butter or suet, it is not acted upon in the mouth, nor in the stomach; but in the small intestines it is broken up by the bile into pieces so small that they can squeeze through the sides of the intestines into some little tubes, called lacteals, along which the fat particles pass, and at last enter the blood. In the blood the fat combines with oxygen, producing heat and force. Some of the fat, however, is not used up in this way, but is deposited round the kidneys or in a layer just under the skin, forming a *store* of food material which can be used by the body if at any time it should be in danger of starvation.

It is found too that fat *aids the digestion* of other foods, though the exact manner in which it does this is not well understood.

We see now the reason why dwellers in cold countries like Greenland and Siberia are so fond of oily and fatty foods, preferring the blubber of the whale or the fat of the seal to all other delicacies. It is because they find in these foods an excellent source of warmth, by which they are enabled to withstand the terrible cold of the long winters of those icy regions. Although oil and butter are also used in hot climates, they are not used so exclusively as in cold countries.

**Functions of Amyloids.**—Under the name of amyloids we include all foods which are chiefly composed either of sugar or of starch. In both sugar and starch the hydrogen and oxygen exist in the proportions in which they combine to form water—namely, two parts

of hydrogen to one part of oxygen. For instance, in the smallest possible piece (molecule) of cane sugar there are 22 atoms of hydrogen and 11 atoms of oxygen; in a molecule of starch there are 10 atoms of hydrogen to 5 atoms of oxygen (see table, page 205). Thus in the amyloids, only carbon is left to combine with fresh oxygen; the hydrogen is already burned up—that is, combined with oxygen.

For this reason amyloids cannot produce so much heat in the body as fatty foods, which contain both carbon and hydrogen ready to combine with more oxygen.

*Digestion of saccharine foods.* Sugar is dissolved in the mouth by the saliva, and it is absorbed by the sides of the stomach, passing directly into the blood,

*Digestion of starchy foods.* Before being taken as food, starch requires to be *boiled*, which causes the grains to swell until they burst. In the mouth a remarkable change takes place: the saliva, acting chemically upon the starch, converts as much of it as it can into grape sugar, which is absorbed as before in the stomach. The remainder of the starch passes through the stomach unchanged, but is also converted into sugar in the small intestines by the action of the pancreatic juice, a fluid formed by an organ called the pancreas. Thus all the starch we eat as food is changed into sugar: this change is absolutely necessary, for starch will not dissolve, and consequently could not pass through the sides of the stomach and intestines to enter the blood. In the blood this sugar made from starch becomes completely burned up or oxidized by its carbon combining with oxygen; and thus, like fat, the amyloids *help to produce and main-*

*tain the heat of the body*, though not in so great a degree. But it is also found that by the action of the liver the sugar in the blood can be changed, first into a substance called "liver-starch," and then into fat: this explains why sugary and starchy foods are so fattening, since we see that within the body they are in part converted into fat.

It is found that about *half a pound* of carbon is required daily in the food of an adult. Now we can obtain this quantity either from three-quarters of a pound of fat or from a pound and a half of sugar. Usually, however, we obtain the necessary eight ounces of carbon partly from one food and partly from another.

## XII.—FUNCTIONS OF MINERAL FOODS.

• **Water as a Mineral Food.**—Of the various foods obtained from the mineral kingdom, perhaps the most important is WATER. Some persons do not consider water to be a food at all; but this is certainly a mistake, since water is taken into the body and absorbed just like any other food. It forms more than three-quarters of the weight of the blood, and every organ of the body is watery and moist. Moreover, a person entirely deprived of water dies in about a week; while if water be frequently taken, it is possible to go without any other nourishment for a much longer period.

Water has various functions to perform in the human body. In the first place, it acts as a *carrier*. The fluid we call blood, which carries the food about to all parts of the body, and removes the waste matter

from every organ, is more than three parts water: more exactly, in 100 lbs. of blood there are 79 lbs. of water.

Secondly, water helps to *dissolve* the food, and so enables it to enter the blood. Sugar, for instance, is dissolved by the watery part of the saliva. Thirdly, water dissolves the used-up materials of the body, and washes them out of the body. In some form or other, every adult takes about four pints (equal to 5 lbs. weight) of water daily, and loses daily the same quantity. The organs which remove water from the body are the kidneys, 50 oz. daily; the skin, 18 oz. (as perspiration); and the lungs, 9 oz. (as water-vapour in the breath).

From the table on page 143, it will be seen that fully two-thirds of the weight of the body of an ordinary adult is water; truly an extraordinary proportion!

**Oxygen.**—Oxygen is another substance which we seldom think of as a *food*. It is a gaseous element, forming one-fifth of the atmosphere, and is most remarkable for its power of aiding the combustion or burning of other substances. About a pound and a half of oxygen gas enters the blood every day through the lungs. In the blood, the function of the oxygen is to *produce and maintain the heat of the body*, by combining with the carbon and hydrogen in the food.

**Common Salt.**—About half a pound of common salt is contained in the body, small quantities of it being found in every organ and in the blood. As a condiment, salt acts by increasing the flow of the saliva; the sodium in it also helps to form bile; while the chlorine aids in making the hydrochloric

acid which gives acidity to the gastric juice. Salt is essential to life. Some people think they eat no salt, but there are small quantities of it in several of the foods which they take.

**Phosphate of Lime** is an important mineral substance, whose function is to give hardness and solidity to the bones. When the bones of little children become bent or diseased, it is generally from want of proper food containing sufficient phosphate of lime. We get this substance mainly in bread, but milk and many other foods contain small quantities of it.

**Potash Salts** are very useful in purifying the blood; they are found in fresh vegetables and in most seeds.

**Iron.**—Although there are only about 65 grains weight of iron in the body, yet it seems to be necessary to health: the *red* colouring matter of the blood contains iron. Most foods contain a little iron; for instance, 6,000 gallons of milk contain about one pound of this metal.





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